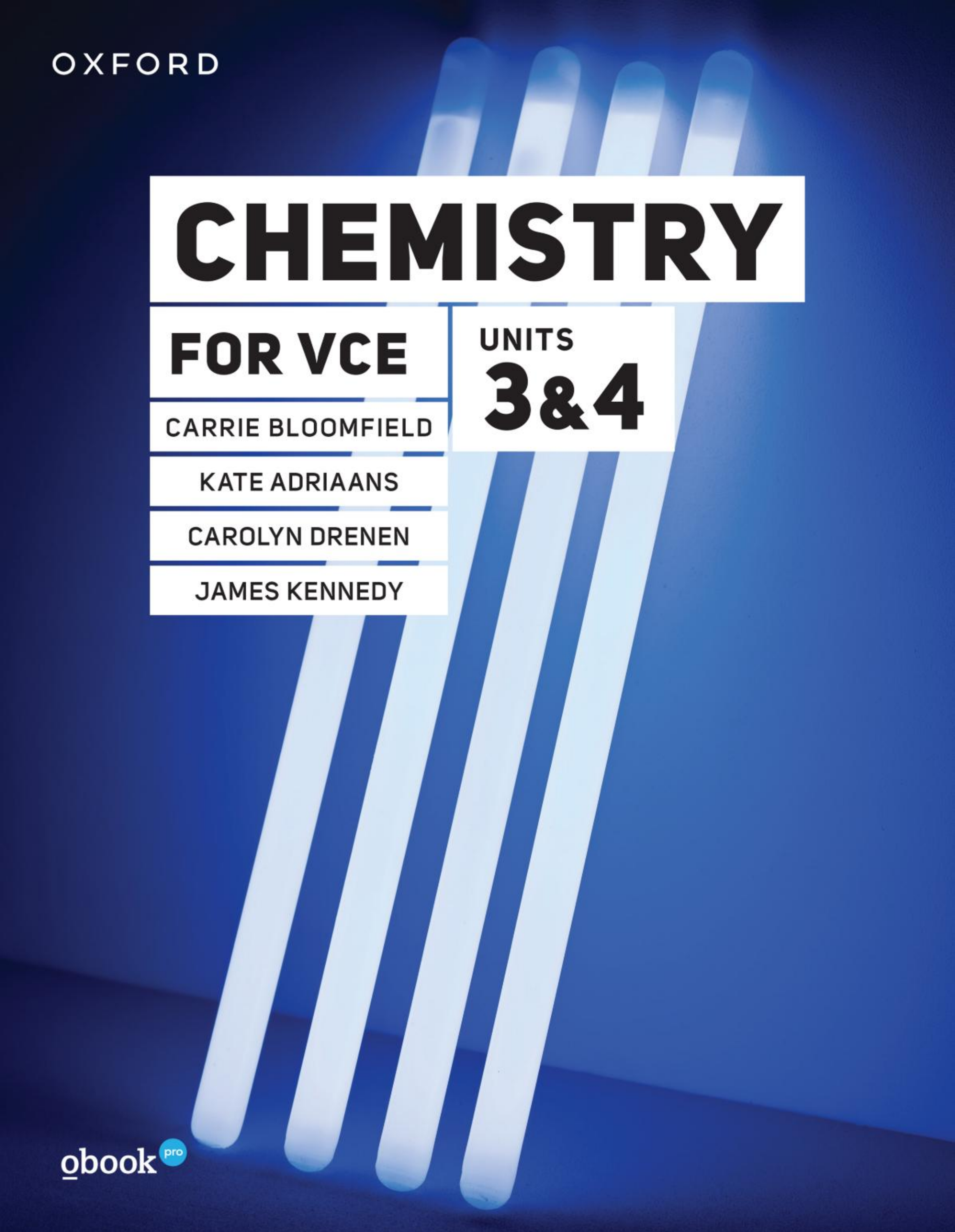

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CHEMISTRY

FOR VCE

UNITS 3&4

CARRIE BLOOMFIELD
KATE ADRIAANS
CAROLYN DRENEN
JAMES KENNEDY

obook pro

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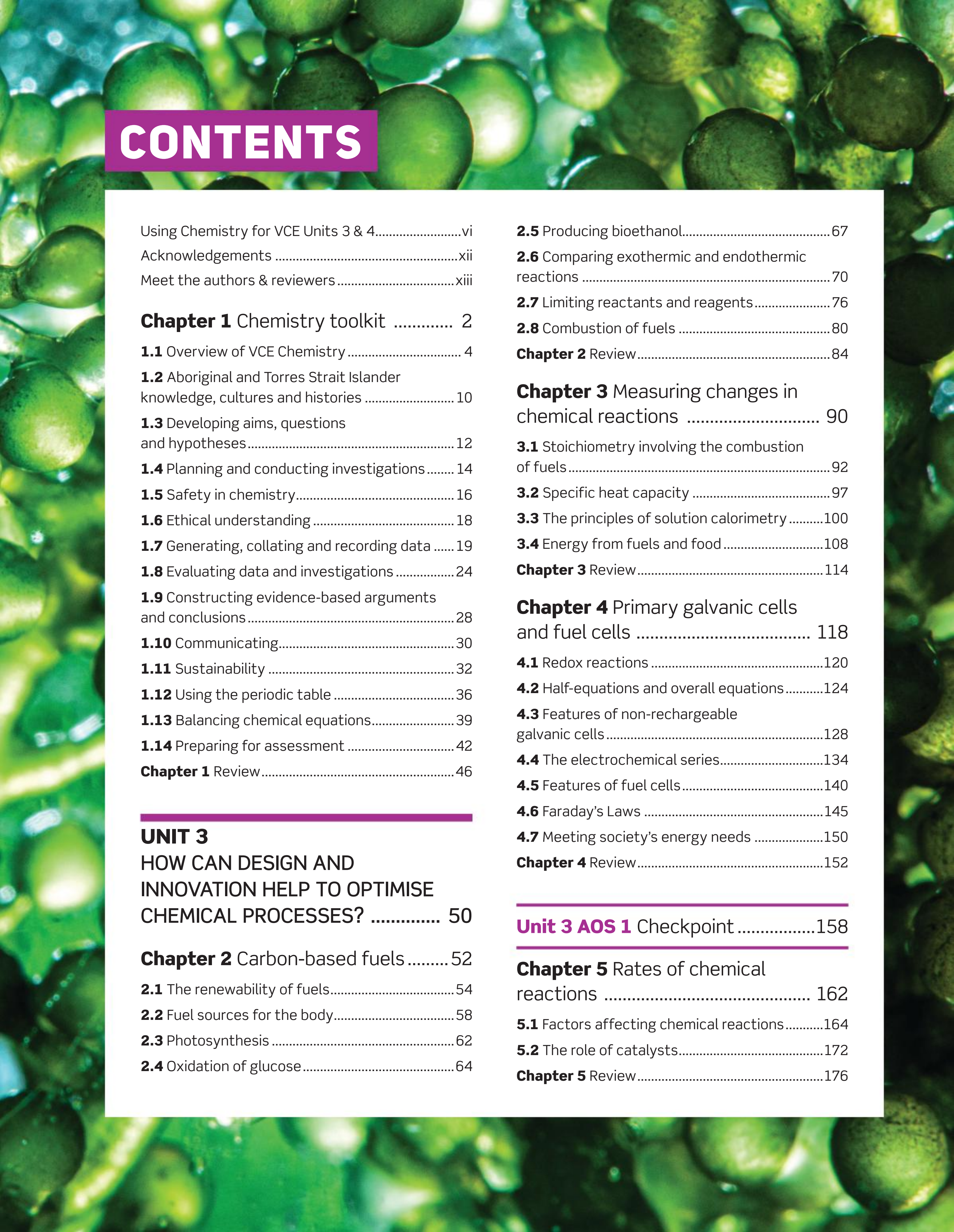


CONTENTS

Using Chemistry for VCE Units 3 & 4

Key features of the Student Book

- This Student Book combines complete coverage of the VCAA Chemistry Study Design 2023–2027 with clear and engaging design.
- Each print Student Book comes with complete access to all the digital resources available on Student obook pro.

Chemistry toolkit
The Student Book begins with a stand-alone reference chapter that includes:

- assessment advice and structured examples
- a step-by-step guide to preparing for your exam
- methods for presenting and analysing chemical data.

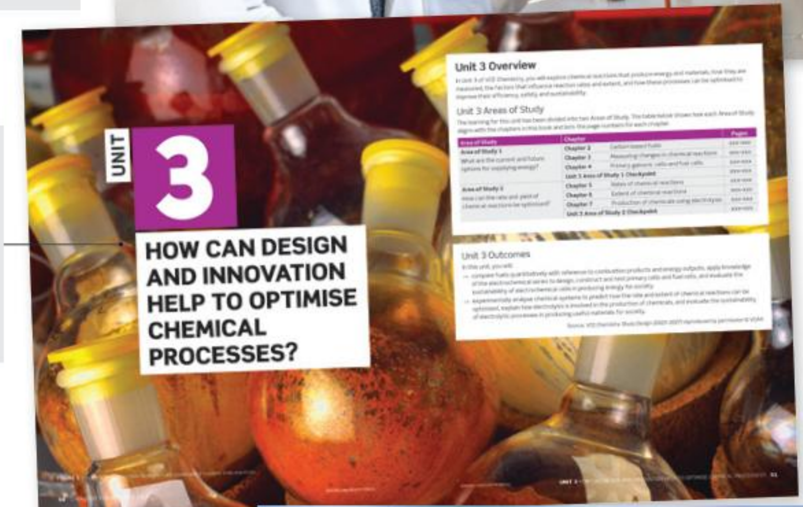

CHAPTER 1 Chemistry toolkit

KEY SCIENCE SKILLS

GROUNDWORK

Unit openers
Each unit begins with a unit opener that includes:

- an **overview of Topics** in the Unit
- **Unit objectives** from the syllabus.


UNIT 3

HOW CAN DESIGN AND INNOVATION HELP TO OPTIMISE CHEMICAL PROCESSES?

Unit 3 Overview

Unit 3 Areas of Study

Unit 3 Outcomes

Groundwork
Questions for students to test their prior knowledge before they begin the chapter. Digital worksheets are available to help refresh their understanding.

Chapter openers
Each chapter begins with a chapter opener that includes key knowledge from the study design.

Practicals
A list of practicals within the chapter that allow students to apply their knowledge to real-life activities.


CHAPTER 12 Medicinal chemistry

KEY KNOWLEDGE

GROUNDWORK

PRACTICALS

12.1

12.3

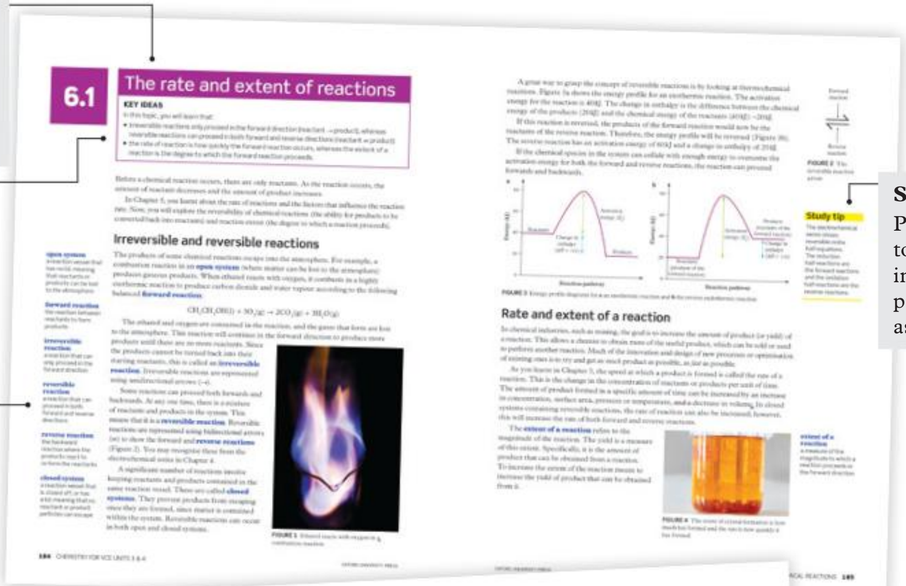
12.4

12.5

Topic-based approach
Content is presented in clearly structured topics. Each topic is labelled and numbered to help navigation.

Key ideas
Concept statements outlining the core content that students should take away from the topic.

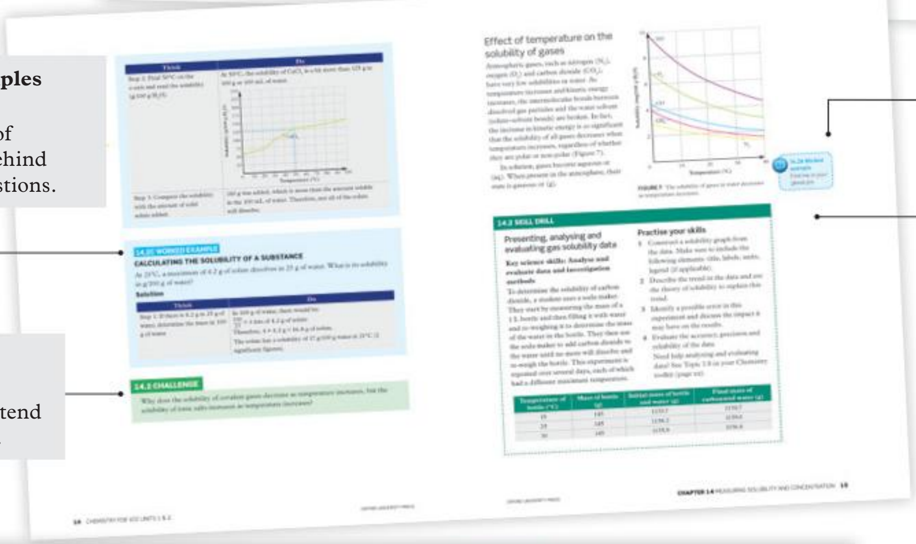
Margin glossary
Definitions to all key terms.



Study tip
Practical advice to help students improve their performance in assessment tasks.

Worked examples
Step-by-step walkthroughs of the thinking behind answering questions.

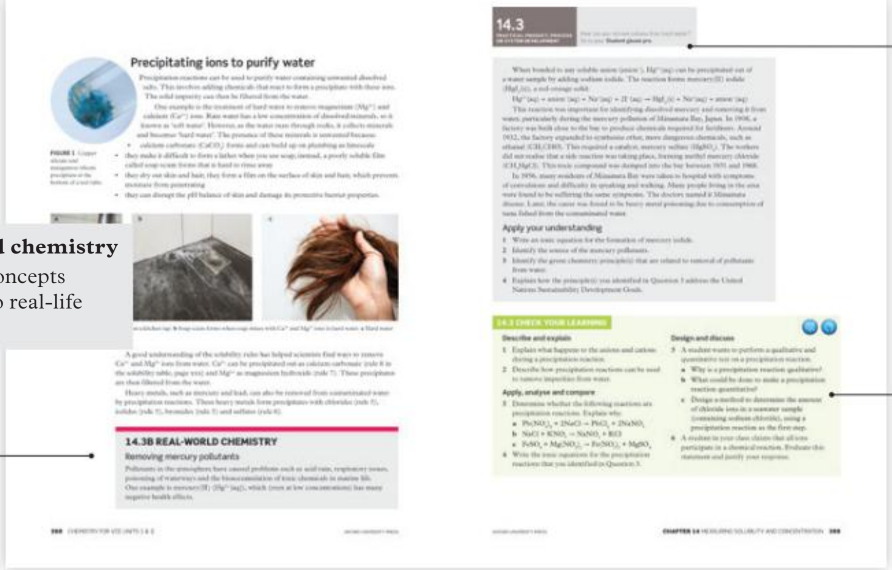
Challenges
More difficult questions to extend understanding.



Digital only features
Dashed boxes for digital only features, accessible via the obook pro.

Skill drills
Students can practise their key science skills in context.

Real-world chemistry
Chemical concepts are linked to real-life examples.

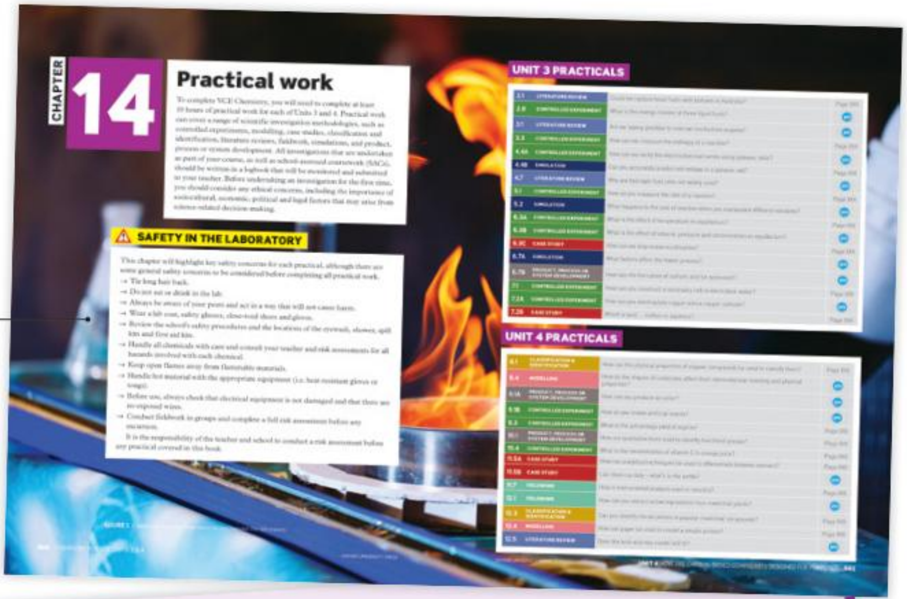


Practical links
Direct links to practicals.

Check your learning
Activity boxes with questions and tasks organised using cognitive verbs according to Bloom's taxonomy.

Practical manual

Each chapter contains a range of practicals that cover all eight scientific investigation methodologies. Practicals are available in the student book and via the obook pro.



CHAPTER 14 Practical work

SAFETY IN THE LABORATORY

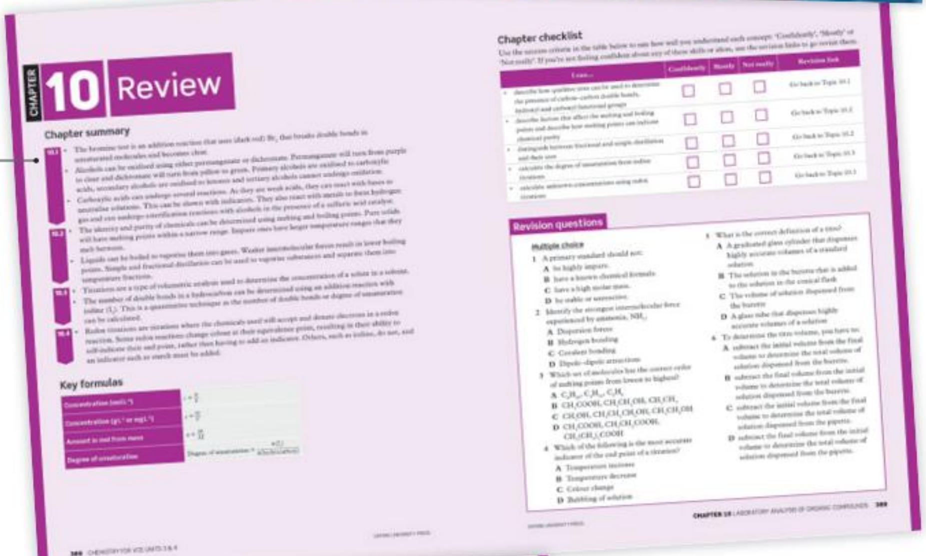
UNIT 3 PRACTICALS

UNIT 4 PRACTICALS

Chapter reviews

Each chapter review includes:

- a summary of key learning in each chapter
- key formulas relevant to the chapter
- a chapter checklist for students to evaluate their understanding of the key knowledge
- revision questions written to target assessment through multiple-choice and short-answer questions.



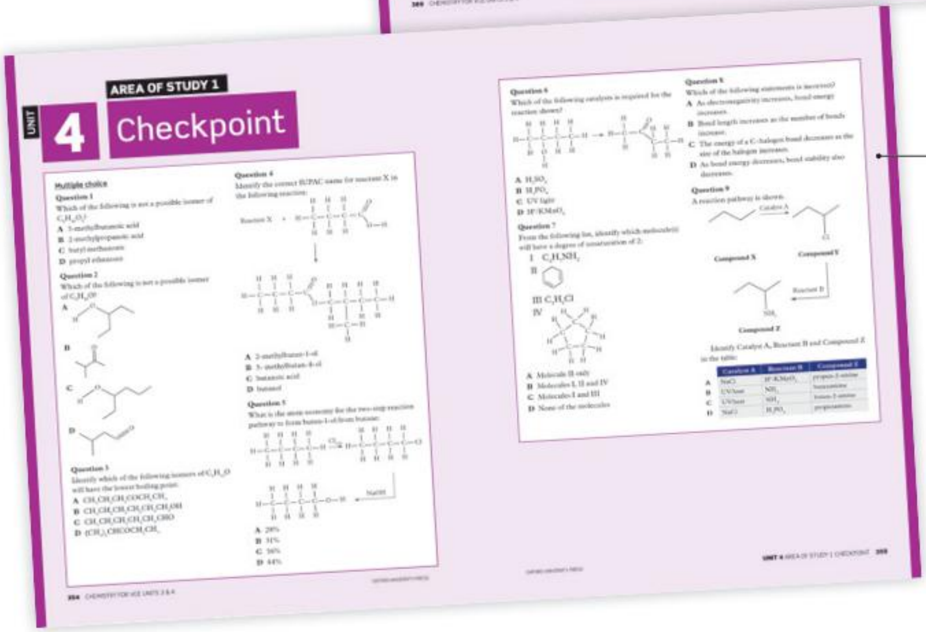
CHAPTER 10 Review

Chapter summary

Key formulas

Chapter checklist

Revision questions



UNIT 4 AREA OF STUDY 1 Checkpoint

Checkpoints

Exam-style revision questions for each Area of Study, including multiple-choice and short-answer questions.

Online only chapter

Guidance for Unit 4 Outcome 3 Student-designed investigation available via the obook pro.

CHAPTER 13 Student-designed investigation

KEY KNOWLEDGE

GROUNDWORK

In Chapter 13, you will learn about how you can prepare for your student-designed investigation, required for Unit 4 Area of Study 3.

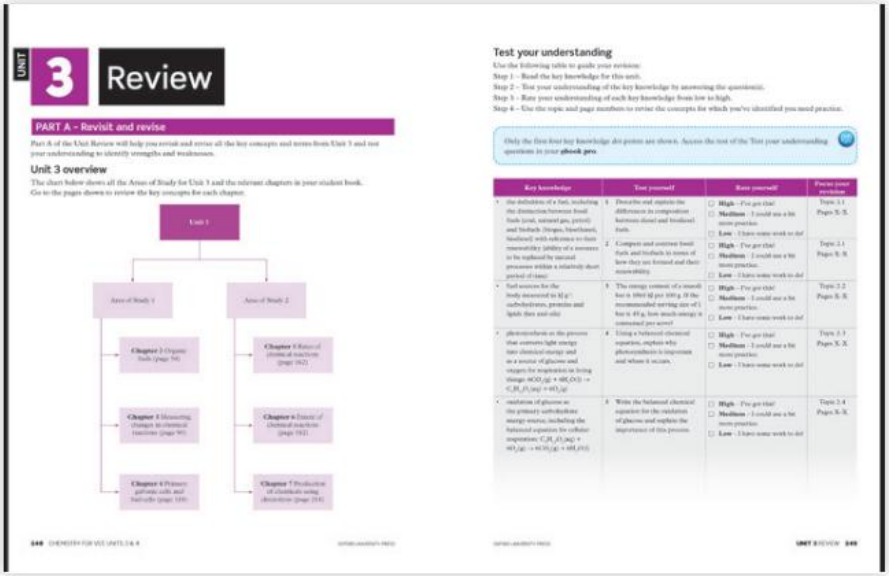
This chapter will build on skills you have already learnt in Units 1 and 2. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

13A Explain what primary data is.

13B Contrast independent, dependent and controlled variables.

13C Identify the different graph types you can use to display a relationship between two variables.

Access this chapter via your Student obook pro



UNIT 3 Review

PART A – Revisit and revise

Unit 3 overview

Test your understanding

Unit reviews

At the end of each unit, students can:

- revisit and revise the content with questions mapped to each key knowledge dot point
- gain tips to succeed in exams and apply their skills in a Think like an examiner activity
- answer practice exam questions.

Digital hotspots

Digital icons or hotspots found throughout the student book link to digital resources accessible via the obook pro.

Video – Watch a video demonstration for the Worked examples and Practicals

Assessment – Access a digital quiz for the topic or chapter

Resource – Access a worksheet or additional resource.

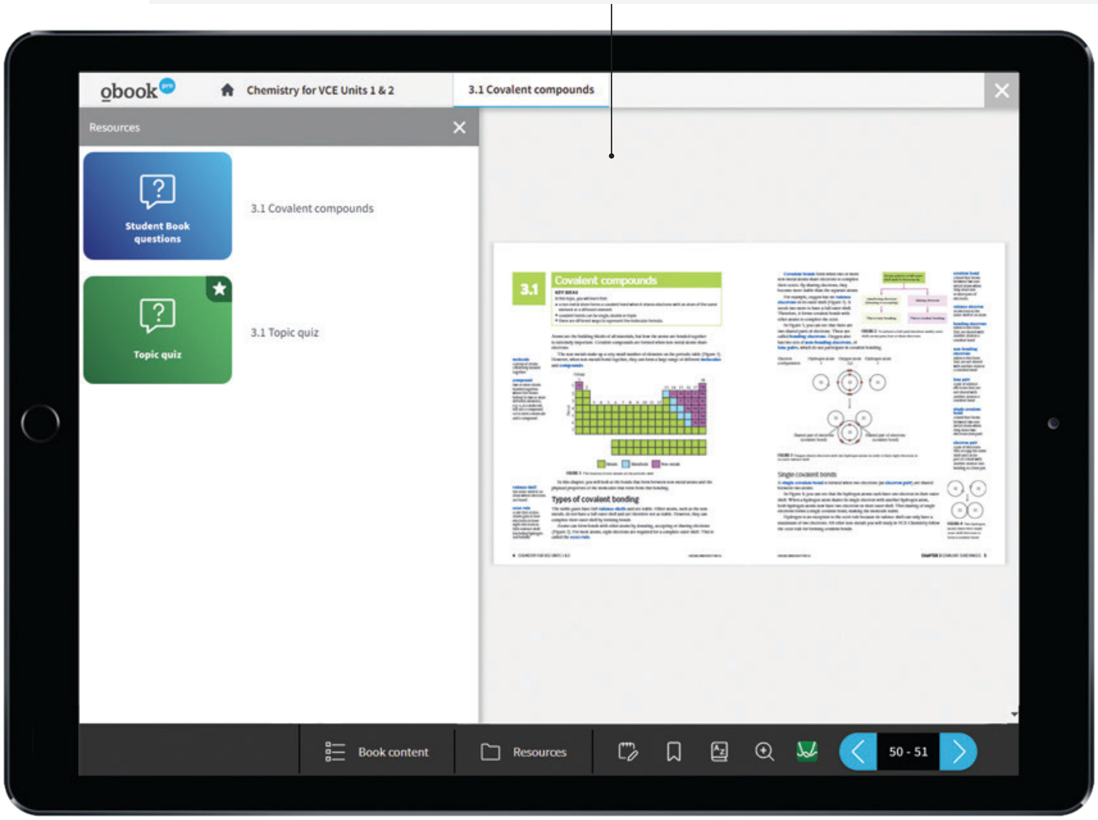
Key features of Student obook pro

- Student obook pro is a completely digital product delivered via Oxford's online learning platform, **Oxford Digital**.
- It offers a complete digital version of the Student Book with interactive note-taking, highlighting and bookmarking functionality, allowing students to revisit points of learning.
- A complete ePDF of the Student Book is also available for download for offline use and read-aloud functionality.

Focus on eLearning

Complete digital version of the Student Book

- This digital version of the Student Book is true to the print version, making it easy to navigate and transition between print and digital.



Interactive quizzes

- Each topic in the Student Book is accompanied by an interactive assessment that can be used to consolidate concepts and skills.
- These interactive quizzes are autocorrecting, with students receiving instant feedback on achievement and progress. Students can also access all their online assessment results to track their own progress and reflect on their learning.
- Each chapter is supported by a multiple choice quiz to give students further practice with exam-style questions.

Benefits for students

- integrated Australian Concise Oxford Dictionary look up feature
- targeted instructional videos for practicals and worked examples
- interactive assessments to consolidate understanding
- integrated Quizlet sets, including real-time online quizzes with live leaderboards
- access to their online assessment results to track their own progress.

Key features of Teacher obook pro

> Teacher obook pro is a completely digital product delivered via **Oxford Digital**.

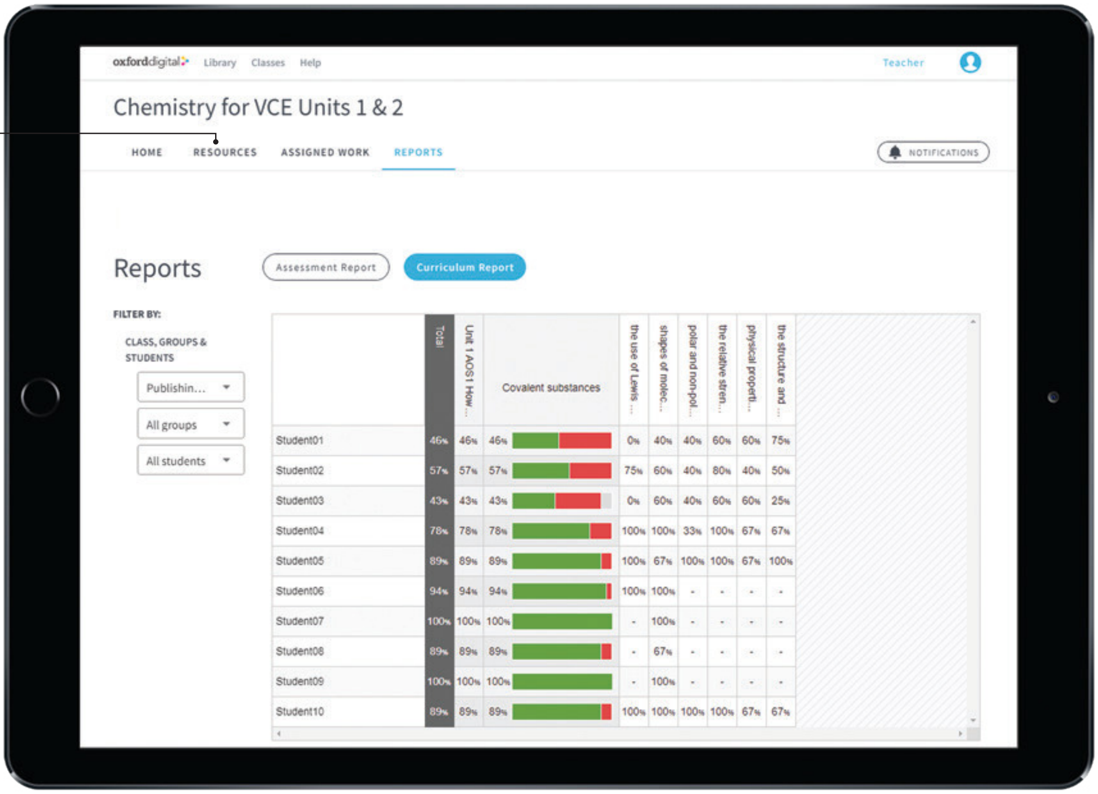
> Each chapter and topic of the Student Book is accompanied by full teaching support. Teacher notes are provided that clearly direct learning pathways throughout each chapter, including ideas for differentiation and practical activities.

> Teachers can use their Teacher obook pro to share notes and easily assign resources or assessments to students, including due dates and email notifications.

Focus on assessment and reporting

Complete teaching support

- Teaching support includes full lesson and assessment planning, ensuring there is more time to focus on students.



Chemistry for VCE Units 1 & 2

Reports

	Total	Unit 1 AOS1 How ...	Covalent substances	the use of Lewis ...	shapes of molec...	polar and non-pol...	the relative stren...	physical properti...	the structure and ...
Student01	46%	46%	46%	0%	40%	40%	60%	60%	75%
Student02	57%	57%	57%	75%	60%	40%	80%	40%	50%
Student03	43%	43%	43%	0%	60%	40%	60%	60%	25%
Student04	78%	78%	78%	100%	100%	33%	100%	67%	67%
Student05	89%	89%	89%	100%	67%	100%	100%	67%	100%
Student06	94%	94%	94%	100%	100%	-	-	-	-
Student07	100%	100%	100%	-	100%	-	-	-	-
Student08	89%	89%	89%	-	67%	-	-	-	-
Student09	100%	100%	100%	-	100%	-	-	-	-
Student10	89%	89%	89%	100%	100%	100%	100%	67%	67%

Additional resources

- Each chapter of the Student Book is accompanied by additional teaching and learning resources to help students progress.

Benefits for teachers

> In addition to online assessment, teachers have access to editable practice exams that are provided in topic 1.14 of the Chemistry toolkit. These exams are formatted like the VCAA Chemistry exam.

> Teachers are provided with laboratory support through experiment answer guidance, laboratory technician notes and risk assessments to ensure safe learning experiences.

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Carolyn Drenen has been teaching VCE Chemistry and Science in secondary schools for the past nine years, and is currently at Lalor North Secondary College. As an ECCN committee member, she has presented workshops at previous VCE Chemistry Conferences since 2015. She also connects with pre-service teachers in her role of university liaison. She authored Oxford University Press's *Chemistry for Queensland Units 3 & 4 Student Workbook*.

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Reviewer: Francis Nguyen

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CHAPTER 1

Chemistry toolkit

KEY SCIENCE SKILLS

- → Develop aims and questions, formulate hypotheses and make predictions
- → Plan and conduct investigations
- → Comply with safety and ethical guidelines
- → Generate, collate and record data
- → Analyse and evaluate data and investigation methods
- → Construct evidence-based arguments and draw conclusions
- → Analyse, evaluate and communicate scientific ideas

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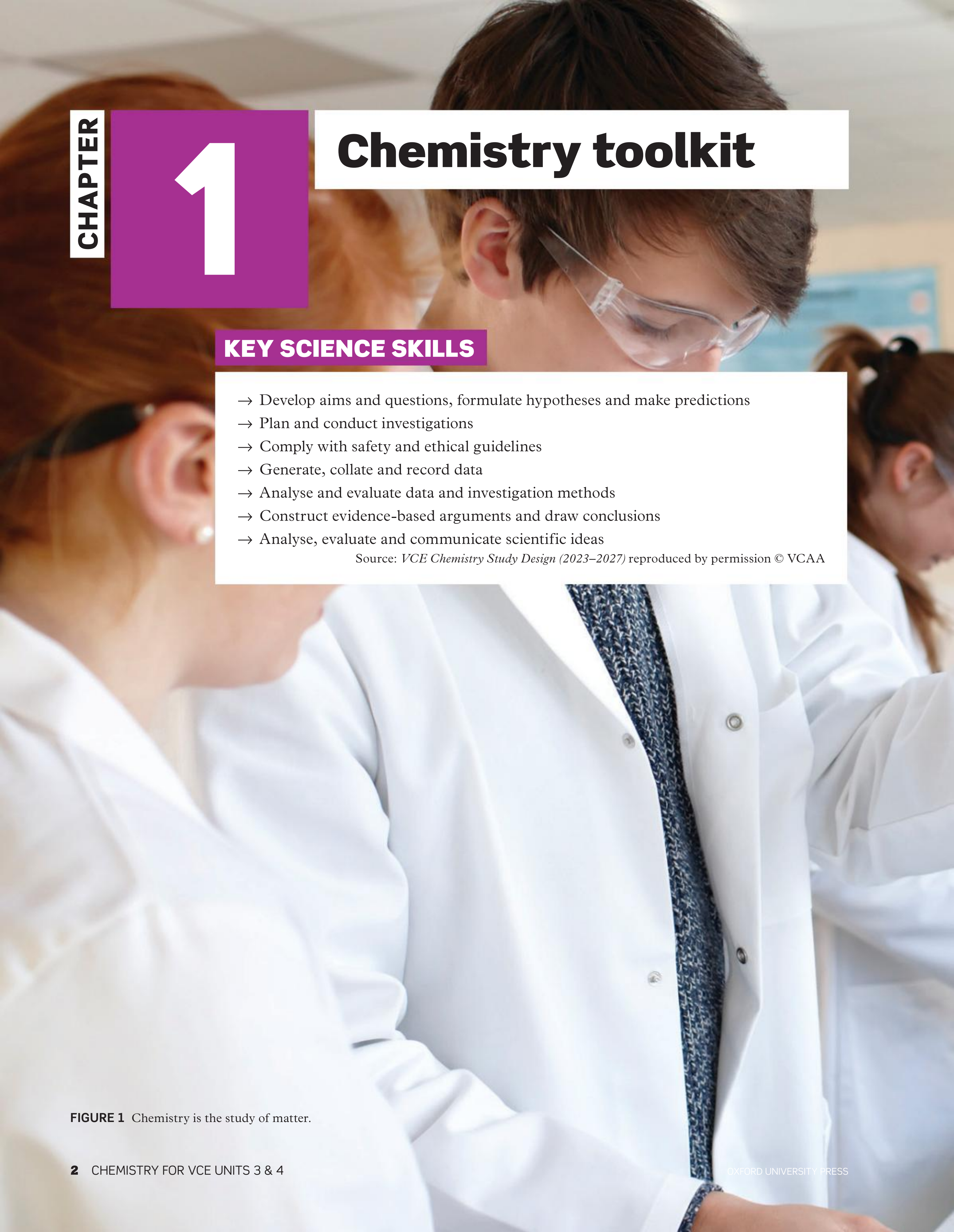

FIGURE 1 Chemistry is the study of matter.

GROUNDWORK

In Chapter 1, you will gain an overview of the VCE Chemistry course and build on your key science skills. You will also learn about the sustainability perspectives that are important for VCE Chemistry and gain skills to succeed in your assessments.

This chapter will build on skills you have already started developing in Chemistry Units 1 and 2. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

1A Explain what a hypothesis is.

1A Groundwork resource
Hypotheses

1B Define qualitative data and quantitative data and give an example of each.

1B Groundwork resource
Qualitative and quantitative data

1C Identify at least three features you should include when presenting a graph or a table.

1C Groundwork resource
Presenting data

1D Contrast a controlled variable, an independent variable and a dependent variable in an experiment or investigation.

1D Groundwork resource
Variables

1.1 Overview of VCE Chemistry

KEY IDEAS

In this topic, you will learn that:

- studying Chemistry can lead to a diverse range of career pathways
- VCE Chemistry is divided into units and areas of study
- the key science skills and their application are important for success in VCE Chemistry.

chemistry the study of matter, including its structure, properties and behaviour

Chemistry is the study of the properties and behaviour of matter. It is often called the 'central science' because key concepts from Chemistry can be linked to other subjects. If you study VCE Biology or VCE Physics, you will see that some key ideas can be linked to what you are learning in VCE Chemistry.

FIGURE 1 Nutritionists use chemistry in their occupation.

Careers in chemistry

Studying Chemistry in VCE can lead to many opportunities when you finish school. Industries that chemists work in include (but are not limited to):

- medicine and health care (e.g. pharmacist, doctor, toxicologist, drug development)
- environment (e.g. environmental chemist, geochemist, water quality scientist)
- manufacturing (e.g. food and drink manufacturing, chemical engineering, materials development)
- science communication (e.g. science writer, teacher, science lecturer, science presenter).

Studying Chemistry in VCE also gives you a wide variety of transferrable skills, which are skills that you can use across many different disciplines. These skills include:

- problem-solving
- teamwork
- critical thinking
- communication
- research.

Structure of the VCE Chemistry course

Chemistry is one of the five science courses offered in VCE. When you study VCE Chemistry, you have an opportunity to engage in a range of inquiry tasks and develop key science skills. You will develop an understanding of the elements, the composition and behaviour of matter and the processes involved when matter reacts to produce useful materials in sustainable ways.

VCE Chemistry consists of four units (see Figure 2). Each unit is divided into Areas of Study. The Areas of Study that you will explore in Units 3 and 4 are outlined in Table 1.

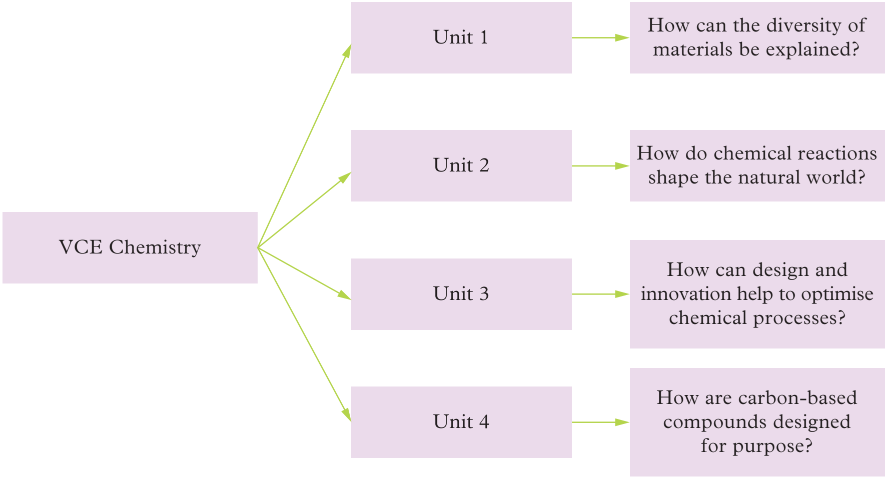

FIGURE 2 The structure of the VCE Chemistry course

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TABLE 1 Areas of Study for Units 3 and 4 Chemistry

<table>
<tr><th colspan="2">Unit 3 How can design and innovation help to optimise chemical processes?</th></tr>
<tr><th>Area of Study</th><th>Description</th></tr>
<tr><td>1. What are the current and future options for supplying energy?</td><td>In this Area of Study, you will learn about:
• fossil fuels and biofuels as energy sources for society
• carbohydrates, proteins and lipids, and fuel sources for the body
• the quantification of energy and gases produced in combustion reactions
• sustainable production of energy from chemicals to meet the needs of society while minimising negative impacts on the environment.</td></tr>
<tr><td>2. How can the rate and yield of chemical reactions be optimised?</td><td>In this Area of Study, you will learn about:
• factors that affect the rate and yield of equilibrium
• electrolytic reactions involved in producing important materials for society
• the qualitative analysis of reactants and products in chemical reactions through the application of Le Chatelier’s principle
• the quantification of reactants and products using equilibrium expressions, reaction quotients and Faraday’s Laws
• sustainability of different options for producing useful materials for society.</td></tr>
<tr><th colspan="2">Unit 4 How are carbon-based compounds designed for purpose?</th></tr>
<tr><th>Area of Study</th><th>Description</th></tr>
<tr><td>1. How are organic compounds categorised and synthesised?</td><td>In this Area of Study, you will learn about:
• the structure, naming, properties and reactions of organic compounds
• chemical reactions associated with the metabolism of food
• sustainable production of synthetic organic compounds for use in society.</td></tr>
<tr><td>2. How are organic compounds analysed and used?</td><td>In this Area of Study, you will learn about:
• laboratory and instrumental analyses of organic compounds
• function of some organic compounds as medicines.</td></tr>
<tr><td>3. How is scientific inquiry used to investigate the sustainable production of energy and/or materials?</td><td>In this Area of Study, you will conduct an investigation that you have designed or adapted yourself. The investigation will relate to the production of energy and/or chemicals and/or the analysis or synthesis of organic compounds.
To do this, you will write your own research question, and then conduct an experiment that generates primary data for you to answer your question.
You will maintain a logbook for record, assessment and authentication purposes.</td></tr>
</table>

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How you will be assessed

In Units 3 and 4, you can be assessed in several different ways for the Areas of Study. These are outlined in Table 2. The assessment in VCE Chemistry is designed to test you against an Outcome for each Area of Study. The Outcome outlines what you should be able to do by the end of the Area of Study if you have completed it successfully.

TABLE 2 Areas of Study 1 and 2 Outcomes and Assessment tasks

Area of Study	Outcome	Assessment
Unit 3 Area of Study 1: What are the current and future options for supplying energy?	The student should be able to compare fuels quantitatively with reference to combustion products and energy outputs, apply knowledge of the electrochemical series to design, construct and test primary cells and fuel cells, and evaluate the sustainability of electrochemical cells in producing energy for society.	For each Outcome in Areas of Study 1 and 2, you could be asked to complete one of the following tasks: • a comparison and evaluation of chemical concepts, methodologies and methods, and findings from at least two student practical activities • an analysis and evaluation of primary and/or secondary data, including identified assumptions or data limitations, and conclusions • problem-solving, including calculations using chemical concepts and skills applied to real-world contexts • an analysis and evaluation of a chemical innovation, research study, case study, socio-scientific issue, or media communication At least one of the four tasks should include reference to sustainability.
Unit 3 Area of Study 2: How can the rate and yield of chemical reactions be optimised?	The student should be able to experimentally analyse chemical systems to predict how the rate and extent of chemical reactions can be optimised, explain how electrolysis is involved in the production of chemicals, and evaluate the sustainability of electrolytic processes in producing useful materials for society.	
Unit 4 Area of Study 1: How are organic compounds categorised and synthesised?	The student should be able to analyse the general structures and reactions of the major families of compounds, design reaction pathways for organic synthesis, and evaluate the sustainability of the manufacture of organic compounds used in society.	
Unit 4 Area of Study 2: How are organic compounds analysed and used?	The student should be able to apply qualitative and quantitative tests to analyse organic compounds and their structural characteristics, deduce structures of organic compounds using instrumental analysis data, explain how some medicines function, and experimentally analyse how some natural medicines can be extracted and purified.	

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The Area of Study 3 Outcome and Assessment task is slightly different.

TABLE 3 Area of Study 3 Outcome and Assessment task

Area of Study	Outcome	Assessment
Unit 4 Area of Study 3: How is scientific inquiry used to investigate the sustainable production of energy and/or materials?	The student should be able to design and conduct a scientific investigation related to the production of energy and/or chemicals and/or the analysis or synthesis of organic compounds, and present an aim, methodology and method, results, discussion and conclusion in a scientific poster.	a report of a student-designed and student-conducted scientific investigation using a structured scientific poster and logbook entries

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You can use Chapter 13 Student-designed investigation to guide you through Unit 4 Area of Study 3.

Key science skills

The key science skills are applicable to all Areas of Study in Units 1–4 of the VCE Chemistry course. They are especially important for planning and conducting investigations for your assessment tasks.

In addition to key knowledge (which we will cover in Chapters 2–12), VCE Chemistry requires you to develop and apply a range of key science skills. These skills are specified in the VCAA VCE Chemistry Study Design and are listed in Table 4.

Study tip

Make sure to use the Skill drills throughout your Student Book to practise your key science skills.

TABLE 4 Key science skills

Key science skill	VCE Chemistry Units 1–4
Develop aims and questions, formulate hypotheses and make predictions (see Topic 1.3)	• identify, research and construct aims and questions for investigation • identify independent, dependent and controlled variables in controlled experiments • formulate hypotheses to focus investigations • predict possible outcomes of investigations
Plan and conduct investigations (see Topic 1.4)	• determine appropriate investigation methodology: case study; classification and identification; controlled experiment; fieldwork; literature review; modelling; product, process or system development; simulation • design and conduct investigations; select and use methods appropriate to the selected investigation methodology, including consideration of sampling technique and size, equipment and procedures, taking into account potential sources of error and causes of uncertainty; determine the type and amount of qualitative and/or quantitative data to be generated or collated • work independently and collaboratively as appropriate and within identified research constraints, adapting or extending processes as required and recording such modifications in a logbook
Comply with safety and ethical guidelines (see Topics 1.5 and 1.6)	• demonstrate safe laboratory practices when planning and conducting investigations by using risk assessments that are informed by safety data sheets (SDS), and accounting for risks • apply relevant occupational health and safety guidelines while undertaking practical investigations • demonstrate ethical conduct when undertaking and reporting investigations

(continued)

TABLE 4 continued

Key science skill	VCE Chemistry Units 1–4
Generate, collate and record data (see Topic 1.7)	• systematically generate and record primary data, and collate secondary data, appropriate to the investigation, including use of databases and reputable online data sources • record and summarise both qualitative and quantitative data, including use of a logbook as an authentication of generated or collated data • organise and present data in useful and meaningful ways, including schematic diagrams, flow charts, tables, bar charts, line graphs and calibration curves
Analyse and evaluate data and investigation methods (see Topic 1.8)	• process quantitative data using appropriate mathematical relationships and units, including calculations of ratios, percentages, percentage change and mean • use appropriate numbers of significant figures in calculations • plot graphs involving two variables that show linear and non-linear relationships • identify and analyse experimental data qualitatively, handling, where appropriate, concepts of accuracy, precision, repeatability, reproducibility, resolution, and validity of measurements; and errors (random and systematic) • identify outliers, and contradictory, provisional or incomplete data • repeat experiments to evaluate the precision of data • evaluate investigation methods and suggest ways to improve precision, and to reduce the likelihood of errors
Construct evidence-based arguments and draw conclusions (see Topic 1.9)	• distinguish between opinion, anecdote and evidence, and scientific and non-scientific ideas • evaluate data to determine the degree to which the evidence supports the aim of the investigation, and make recommendations, as appropriate, for modifying or extending the investigation • evaluate data to determine the degree to which the evidence supports or refutes the initial prediction or hypothesis • use reasoning to construct scientific arguments, and to draw and justify conclusions consistent with the evidence and relevant to the question under investigation • identify, describe and explain the limitations of conclusions, including identification of further evidence required • discuss the implications of research findings and proposals
Analyse, evaluate and communicate scientific ideas (see Topics 1.10 and 1.11)	• use appropriate chemical terminology, representations and conventions, including standard abbreviations, graphing conventions, algebraic equations, units of measurement and significant figures • discuss relevant chemical information, ideas, concepts, theories and models and the connections between them • analyse and explain how models and theories are used to organise and understand observed phenomena and concepts related to chemistry, identifying limitations of selected models/theories • critically evaluate and interpret a range of scientific and media texts (including journal articles, mass media communications and opinions in the public domain), processes, claims and conclusions related to chemistry by considering the quality of available evidence • apply sustainability concepts (green chemistry principles, development goals and the transition from a linear towards a circular economy) to analyse and evaluate responses to chemistry-based scenarios, case studies, issues and challenges • identify and explain when judgements or decisions associated with chemistry-related issues may be based on sociocultural, economic, political, legal and/or ethical factors and not solely on scientific evidence • use clear, coherent and concise expression to communicate to specific audiences and for specific purposes in appropriate scientific genres, including scientific reports and posters • acknowledge sources of information and assistance, and use standard scientific referencing conventions

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1.1 CHECK YOUR LEARNING

Describe and explain

1 Summarise how your knowledge and skills will be assessed in VCE Chemistry Units 3 & 4.

2 Reflect on the key science skills listed in Table 4 and identify 3–5 skills you need to practise most.

Apply, analyse and compare

3 Follow the link in your obook pro to access the 'Chemistry bullseye poster'. Research two different chemistry-related occupations that interest you and answer the following questions.

- **a** List some of the types of activities that a person does in these occupations.
- **b** Identify the types of education or qualifications that are recommended in order to work in these occupations.
- **c** Describe the qualities or interests a person who is interested in these occupations could have.
- **d** Explain what you think is the most interesting part of these occupations.
- **e** Explain what you think is the most difficult or repetitive part of these occupations.

Design and discuss

4 Discuss the results of your research from Question 3 with another student.

FIGURE 3 Examples of occupations that require knowledge of chemistry are: **a** an anaesthetic technician, **b** a chemical engineer, **c** a veterinary nurse and **d** a patents examiner. You might like to investigate what one of these professionals does as part of Question 3.

1.2 Aboriginal and Torres Strait Islander knowledge, cultures and histories

KEY IDEAS

In this topic, you will learn that:

- Aboriginal and Torres Strait Islander Peoples are Australia's first chemists and apply their unique scientific knowledge about the natural world to find (or refine) suitable food sources, and develop traditional medicines and other useful materials in sustainable ways
- Aboriginal and Torres Strait Islander Peoples come from many diverse nations, each of which have their own unique cultures, histories and knowledges.

Aboriginal and Torres Strait Islander cultures are some of the oldest continuous cultures in the world and have existed on this land for at least 60 000 years. The Traditional Custodians of Australia are people who are of Aboriginal and/or Torres Strait Islander descent, and who identify and are accepted as such by their community.

Aboriginal and/or Torres Strait Islander Peoples people with family heritage from, and membership in, the ethnic groups that lived in Australia before British colonisation

As a collective group, **Aboriginal and/or Torres Strait Islander Peoples** may also be referred to as Indigenous Peoples, First Peoples and/or Traditional Custodians/Owners. Indigenous Peoples may also identify with their particular Community, Nation, Tribe or Clan that they, their family and/or their community recognise.

We say Aboriginal and Torres Strait Islander Peoples (not 'people') because there are many diverse language groups in Australia (representing more than 250 languages), each of whom have their own unique cultures, laws, practices and languages. So, when learning or writing about the cultures and practices of Aboriginal and/or Torres Strait Islander Communities, Nations, Tribes or Groups in VCE Chemistry, it is important to identify and acknowledge the right owners of that knowledge.

FIGURE 1 Aboriginal Language groups and Nations within the borders of Victoria

Aboriginal and Torres Strait Islander knowledges and cultures in VCE Chemistry

The diverse cultures and knowledge systems of Aboriginal and Torres Strait Islander Peoples have been acknowledged in the VCE Chemistry course.

In VCE Chemistry, you will learn how scientific thinking can be enhanced by considering how Aboriginal and Torres Strait Islander Peoples have developed and refined their own knowledge, including the use of plants as medicine and the use and modification of natural materials for useful purposes.

For example, Peoples of the Kulin Nation in Victoria can use the sap of the river red gum (*Eucalyptus camaldulensis*) to seal burns, or the smoke of older manna gum (*Eucalyptus viminalis*) leaves to reduce fever. In Western Australia, the Noongar Peoples use infusions of crushed bracken fern (*Pteridium esculentum*) to relieve sores, or the juice from the same plant to relieve insect bites. In New South Wales, the Bidjigal and Gadigal Peoples use the long leaves of the gymea lily (*Doryanthes excelsa*) to produce string, and its long stem to craft a fishing spear.

FIGURE 2 Aboriginal and Torres Strait Islander Peoples hold specialised knowledge about the plants and animals on Country, such as how to use **a** the leaves of the manna gum (*Eucalyptus viminalis*) in Victoria, **b** the bracken fern in Western Australia or **c** the gymea lily in eastern New South Wales.

This knowledge has been accumulated and refined over thousands of years. It is highly sophisticated and specialised for the areas in which the people live.

Study tip

In Unit 1 Area of Study 3, you might research how Aboriginal and Torres Strait Islander Peoples sustainably modify and process raw materials. It is important to be able to carefully evaluate your resources to make sure you're using appropriate ones. See your teacher for more information on how to do this.

1.2 CHECK YOUR LEARNING

Apply, analyse and compare

1 Identify the Traditional Owners of the land your school or home is built on.

Design and discuss

2 Explain why modern medicine would reference the chemical knowledge of Aboriginal and Torres Strait Islander Peoples to treat ailments and infections.

3 Research one of the following topics.
- The use of plants for traditional medicines in Victoria
- The chemical processes that occur when Aboriginal and Torres Strait Islander Peoples detoxify poisonous items (such as cycad seeds) before eating them
- The traditional use of animal fat and/or plant pigments to make paint

4 Evaluate the resources you used in Question 3. Justify your response.

1.3 Developing aims, questions and hypotheses

KEY IDEAS

In this topic, you will learn that:

+ a research question states the specific problem or issue on which your investigation will be based
+ an aim is a statement of what is to be investigated
+ a hypothesis is a testable statement that should include a prediction about the outcome of an investigation, based on scientific reasoning.

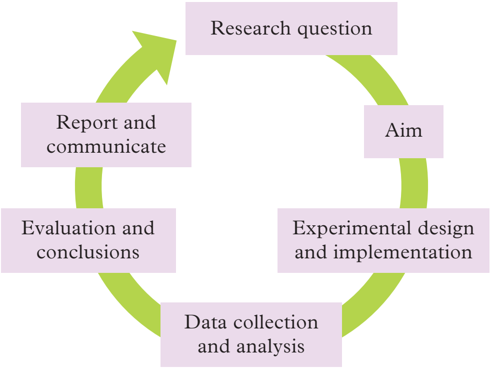

FIGURE 1 The general process of the scientific method. Adapted from: VCAA (2016) *Victorian Certificate of Education Chemistry Advice for Teachers 2016–2020*, page 6, Melbourne, Victorian Curriculum and Assessment Authority.

The **scientific method** is a cyclical process involving a number of steps (see Figure 1):

- developing a research question or aim
- formulating a hypothesis
- conducting experiments to collect, analyse and communicate data (including potential errors and uncertainty)
- developing conclusions
- communicating scientific ideas and understandings.

Developing a research question

In VCE Science, you will need to develop a research question for your student-designed investigation (Unit 4). When writing a research question, you should make sure the question is:

- **clear and specific**: your research question must contain enough detail for your audience to understand what you are investigating
- **concise**: your research question must be expressed in as few words as possible
- **complex**: your research question cannot be answered with a simple ‘yes’ or ‘no’, or with facts that are easily found; it must require enough scope for you to respond with an appropriate experiment with collection and analysis of primary and/or secondary data.

An example of a suitable research question for a scientific inquiry in VCE Chemistry is: Which plant biomass produces an efficient mass per unit volume of essential oil when extracted by steam distillation?

scientific method
a series of steps used to acquire knowledge in science, which involves observation, developing and testing hypotheses; collecting, analysing and communicating results

Writing an aim

An **aim** is a short statement outlining what is to be investigated. The goal of your experiment or inquiry must be precisely stated. An example of an aim for a scientific inquiry in VCE Chemistry is:

To investigate the relationship between plant biomass and the volume of essential oil produced during extraction by steam distillation.

aim
a statement that describes the purpose of an investigation

Independent and dependent variables should also be included where possible, without detail on the method or procedure used. Remember, an **independent variable** is what you change during a controlled experiment. The **dependent variable** is what you measure (it may or may not be affected by the independent variable). A **controlled variable** is one that is not changed or remains constant during an experiment.

Formulating a hypothesis and making a prediction

A **hypothesis** is a testable statement that predicts how the independent variable will affect the dependent variable in an investigation. It should include an explanation for the prediction based on scientific knowledge or reasoning.

A hypothesis cannot be wrong, only supported or negated by the investigation results.

If … then … because …

A good way to make sure you include everything in your hypothesis is to use a structure such as '*If … then … because*' or '*When … the … because*' statement.

Suggested phrases for formulating a hypothesis are listed in Figure 2.

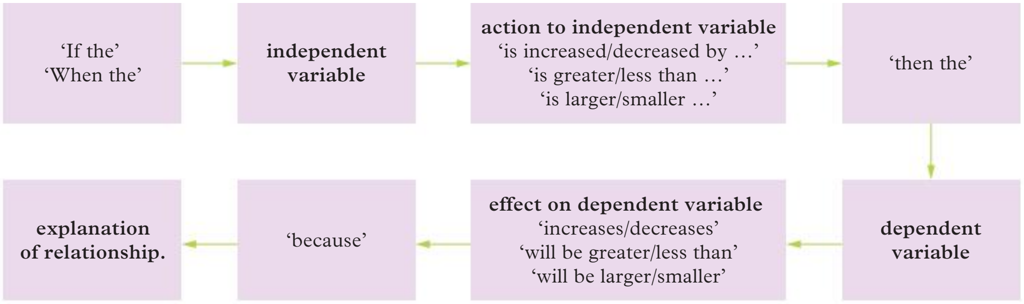

FIGURE 2 Constructing a hypothesis from a research question

An example of a hypothesis for a scientific inquiry in VCE Chemistry is:

When cloves are used, the volume of oil extracted by steam distillation is greater than when eucalyptus leaves are used, because its boiling point is further away from that of the solvent used.

See how to write an aim, research question and hypothesis in Worked example 1.3.

1.3 CHECK YOUR LEARNING

Describe and explain

1 Suggest why the steps of the scientific method, as illustrated in Figure 1, are often shown as a circular process.

Apply, analyse and compare

2 Contrast an aim and a hypothesis.

Design and discuss

3 During studies of acids and bases, a student observed that when she poured lemon juice onto a plant, it shrivelled and died. The student wanted to use this observation as the basis for a practical investigation.

a Design a research question for this practical investigation.

b Write an aim for this practical investigation.

c Formulate a hypothesis for this investigation.

independent variable
the variable that is changed or manipulated during an investigation

dependent variable
the variable that is observed or measured when the independent variable is changed during an investigation

controlled variable
a variable that is not changed or is constant during an investigation

hypothesis
a testable statement that includes a prediction about the outcome of an investigation based on scientific reasoning

Study tip

Make sure you know how to write a hypothesis. This is the kind of skill that is often tested in assessment tasks or exams.

1.3 Worked example
Find me in your obook pro

1.3 Worked example
Video demonstration

1.4 Planning and conducting investigations

KEY IDEAS

In this topic, you will learn that:

- a variety of methodologies can be used for a scientific investigation
- a logbook is required for recording raw data and must be submitted intermittently for assessment purposes.

In VCE Chemistry, you will complete at least 10 hours of practical work per unit.

Scientific investigations are important to the VCE Chemistry course and can be used to collect primary data or data from secondary sources. You can complete an investigation individually, in a small group or with the whole class; however, all the work required for assessment purposes (e.g. the logbook and the poster) must be your own work.

Scientific investigation methodologies

methodology
a system of methods used in a particular area of study or activity

A variety of **methodologies** can be used when planning and conducting scientific investigations. A methodology refers to the approach you take to answer your scientific question. Not all scientific questions should or need to be answered with a controlled experiment. There are other ways to find the answers to your scientific questions or collect data. Some of the methodologies that you will encounter in VCE Chemistry are outlined in Table 1.

TABLE 1 Key scientific investigation methodologies in VCE Chemistry

Investigation methodology	Investigation outline
Case study	A case study involves the investigation of a specific event, an activity or a problem that contains a real or hypothetical situation. Case studies can be historical, involving the analysis of causes and consequences, and/or discussion of knowledge learnt from the situation; a real situation or a role-play of an imagined situation, where plausible recommendations are to be made; or problem-solving, where developing a new design, methodology or method is required.
Classification and identification	Classification refers to the arrangement of phenomena, objects or events into smaller, more manageable groups (e.g. classifying organic compounds into functional groups). Identification is a process of recognising things as belonging to particular sets or possibly being part of a new set.
Controlled experiment	A controlled experiment investigates the relationship between an independent variable and a dependent variable. All other variables are controlled so they don't affect the outcome of the investigation.
Fieldwork	Fieldwork involves going to a specific location to investigate a phenomenon or problem that is unique to that site. When completing fieldwork, you should record site-specific data in your logbook, such as observations and data collected about the conditions of the environment (e.g. air temperature, pH of the water).
Literature review	A literature review involves researching, gathering and interpreting secondary sources of data (also called 'literature' by academics) to answer a research question. This may be used to answer questions that have already been asked by other scientists, or to develop a background understanding before you start your own investigation.
Modelling	A model physically, conceptually or mathematically simulates a concept to assist understanding (e.g. modelling molecule structures, which are too small for us to observe in the classroom).

(continued)

TABLE 1 continued

Investigation methodology	Investigation outline
Product, process or system development	Product, process or system development involves designing an object, process or system to meet a need. This should link scientific knowledge to technological developments.
Simulation	Simulations involve using an existing model to investigate a scientific phenomenon by manipulating variables in the simulation. Simulations are useful when variables cannot be manipulated in real life.

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Logbook

You will need a logbook in hard copy or a digital form to keep primary and secondary data records from all types of scientific investigations undertaken in Units 1–4. You will need to submit your logbook to your teacher for some assessment tasks so they can check that the work you have submitted is your own.

Each entry must include the date and information you add in chronological order, and an acknowledgement of all secondary resources, as well as all expert advice and teacher assistance received.

Your logbook could also include the following information as you complete your investigation:

- planning notes for experiments
- a description of the activities you have carried out
- the results/data from guided activities or investigations (including outliers and/or risk identification and management)
- personal reflections made during or at the conclusion of demonstrations, activities or investigations
- any links to spreadsheet calculations or other digital records and presentation you might use
- any notes and electronic (or other images) taken on excursions, or any database extracts
- notes of any additional work completed outside of class time.

A well-organised logbook that contains all of this information will make it a lot easier to complete your scientific poster in Unit 4 Area of Study 3.

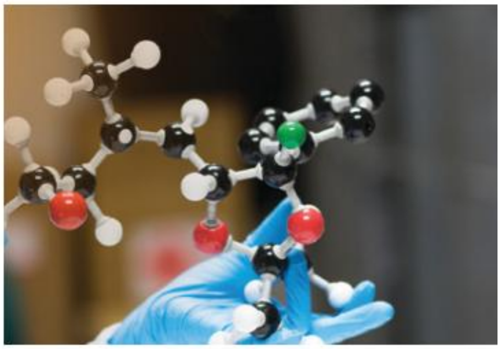


FIGURE 1 Modelling can help us to investigate scientific phenomenon if we are unable to observe the real thing in the lab.

FIGURE 2 Fieldwork allows us to investigate phenomena at a specific location. For example, this scientist is collecting a medicinal plant to extract active organic compounds from.

1.4 CHECK YOUR LEARNING

Describe and explain

1 Describe the advantages of keeping a logbook for investigations in VCE Chemistry.

2 Explain the differences between primary and secondary data.

Apply, analyse and compare

3 Compare:

a a controlled experiment and fieldwork

b a case study and a literature review

c modelling and a simulation.

Design and discuss

4 Design a potential research question for one of the scientific investigation methodologies.

1.5

Safety in chemistry

KEY IDEAS

In this topic, you will learn that:

- safety must be your priority when performing scientific investigations
- a risk assessment is an organised way of identifying hazards and/or risk factors, and to implement controls for prevention when performing scientific investigations.

Chemists often work with dangerous goods and hazardous substances. You might encounter some of these dangerous substances when conducting practical work in VCE Chemistry.

Some hazards are clearly labelled for us. (See Figure 1 for examples of chemical hazard warning symbols.) But other hazards still need to be avoided to keep everyone safe. We can do this by following basic laboratory safety rules and conducting risk assessments.

Rules for laboratory safety

- **Always wait for teacher permission** before entering the laboratory, collecting materials and/or equipment and beginning your activities. Teacher and laboratory technician instructions must also be obeyed for OHS (occupational health and safety) reasons.
- **Always wear the correct personal protective equipment appropriately**. Laboratory coats should be buttoned up to protect your clothes. Laboratory glasses or goggles should be worn on the face for eye protection. Laboratory gloves should be worn whenever possible to protect skin from heat, cold and chemicals. Closed-toe shoes should be worn for foot protection and long hair should be tied back behind the neckline.
- **Always behave sensibly**. Treat the laboratory as a proper workplace environment. Be courteous and respectful to your peers, teachers and other staff who may enter.
- **Never eat or drink in the laboratory**. You don't know what substances have been in the laboratory previously, and ingestion of hazardous materials could be harmful.
- **Report all accidents immediately to the teacher in charge**, including spills, leaks, breakages, and faulty and/or damaged equipment, to minimise escalation of potential hazards.
- **Turn off all equipment and/or pack it away in a careful and safe manner** (for future use by others). Place waste materials and/or chemicals in appropriate containers for disposal and leave your workplace area clean, dry and tidy.

FIGURE 1 The Globally Harmonised System of Classification and Labelling of Chemicals (GHS) is used in Australia.

Risk assessments

A **risk assessment** is an organised way to identify:

- any risk factors and/or hazards that might occur during an investigation
- controls or actions to prevent accidents, injuries or harm to you (or anyone else who might be in the laboratory).

As part of VCE Chemistry, you may need to write and submit a risk assessment to accompany investigations that you have designed yourself.

FIGURE 2 When conducting experiments, you should wear a lab coat, gloves, safety glasses and closed-toe shoes, and tie back long hair.

Risk assessments are written for all workplace environments, including the laboratory, to meet model work health and safety acts, legislation and codes of practice.

risk assessment
a process of evaluating the potential risks that may be involved in an activity, e.g. performing an experiment

Risk assessments can be written in a variety of formats and some schools may use external programs to generate them. An example of a risk assessment for a laboratory-based activity (with annotations) is provided in your obook pro. A blank risk assessment template is also provided.

It is important to consider the information in each section of the risk assessment before assigning a final risk judgment. For example, to reduce the risk of using a hazardous substance, it could be:

- diluted
- dispensed from a dropper bottle
- used by a person who is wearing their personal protective equipment appropriately
- used when following standard laboratory rules.

It is also important to implement and follow the control measures outlined for all potential hazards, physical and/or chemical, in the risk assessment, to ensure the safety of everyone involved in the activity.

1.5 Annotated risk assessment
Find me in your obook pro

1.5 Risk assessment template
Find me in your obook pro

1.5 CHECK YOUR LEARNING

Describe and explain

1 Explain why it is important to comply with safe work practices when planning and conducting laboratory-based activities.

2 Describe the process of how to make a final risk judgment of a laboratory-based activity process or procedure, using an example of a practical activity or teacher demonstration from Units 3 or 4.

Apply, analyse and compare

3 Compare the similarities and differences between the safety data sheets of one GHS substance classified as hazardous and another one classified as non-hazardous.

Design and discuss

4 An experiment is to be conducted to investigate the factors that affect reaction rates, using vitamin C tablets, assorted glassware, a Bunsen burner and water. Design a risk assessment for this experiment.

1.6 Ethical understanding

KEY IDEAS

In this topic, you will learn that:

+ ethics are applied when performing all types of scientific investigations in VCE Chemistry.

ethics
moral principles that govern a person's behaviour or how an activity is conducted

Study tip

When collecting or reporting the results of scientific investigations, be careful to minimise data errors, be honest (do not mislead, or change or make up data) and avoid bias from unexpected data or results.

Ethical understanding is applied across Units 1–4 of VCE Chemistry, especially when performing scientific investigations, analysing data collected from primary and secondary sources, and identifying and investigating any issues relating to the application of scientific knowledge in society.

In VCE Chemistry, **ethics** can include:

- thinking about the impact of scientific investigations (including your own) on living and non-living things and the environment
- being honest when recording data and presenting information from your own investigations, and acknowledging when you are using someone else's ideas or data
- developing an opinion or argument for or against science-related ethical issues based on your understanding of ethical concepts (considering current and future needs) and scientific knowledge
- identifying factors that might influence decision making in science (such as personal values or economic, political or legal factors).

1.6 CHALLENGE

Ethical issues in chemistry

Modern ethical issues in chemistry include: manufacture and use of 'designer' polymers and pharmaceuticals; sustainable consumption of materials and goods; price and/or safety of drugs made by pharmaceutical companies.

Research one of these issues and answer the following questions.

1 What known facts are presented?
2 What ethical concerns do they raise?
3 Which of the known facts are *relevant* to resolving the ethical concerns?
4 What additional facts might be relevant to the case?
5 How might the additional facts affect what is ethically at stake?
6 What options are available, and which seem the best from an ethical viewpoint?

FIGURE 1 Products developed by pharmaceutical companies can be expensive. Is this ethical?

1.6 CHECK YOUR LEARNING

Describe and explain

1 Explain why it is important to understand and apply ethics in VCE Chemistry.

Apply, analyse and compare

2 Download a copy of the *Global Chemists' Code of Ethics*, from your obook pro. Compare this document with the key knowledge outcomes from the VCE Chemistry Study Design, giving reasons for how well they align.

Design and discuss

3 Design an A4-sized report or brochure of the arguments for or against (based on collected evidence) one of the ethical issues in Challenge 1.6. Identify and justify your own position on the issue.

1.7

Generating, collating and recording data

KEY IDEAS

In this topic, you will learn that:

- data generated from scientific investigations is important evidence to support the trends, patterns and/or relationships between variables
- raw data can be qualitative or quantitative, discrete or continuous
- raw data from scientific investigations needs to be presented clearly so that it can be easily understood by your audience.

In VCE Chemistry and other sciences, investigations are important for developing explanations, supported by evidence (or data), to explain natural phenomena and events. Raw data generated from scientific investigations and other evidence recorded in your logbook must be presented in an appropriate format to illustrate any trends, patterns and/or relationships between the independent, dependent and controlled variables.

Types of data

From previous studies in statistics or mathematics, you know that there are many types of raw (or unprocessed) data that can be collected. Studies in VCE Chemistry involve the following raw data types.

- **Qualitative data** can be described in words, phrases or categories (e.g. classifying types of substances or describing investigation observations in dot points or sentences).
- **Quantitative data** can be described with numbers, quantities or other numerical values, often collected during scientific investigations. Quantitative data can be further classified as:
  - **discrete data** – set numerical values that are not related (e.g. comparing the energy content of different foods or fuels)
  - **continuous data** – numerical values within a specified range that are usually generated when measurement is involved (e.g. temperature, mass, length, voltage, volume, time, pH).

qualitative data
data that is not numerical, can be text, images or audio

quantitative data
data that can be counted or measured and is expressed as numbers

discrete data
data or information that can only be certain numerical values

continuous data
data or information that can be any numerical value

collate
collect and combine texts, information or data

Collating and recording data

All raw data from scientific investigations, either qualitative or quantitative, should be **collated** and presented in a table, which has a title and labelled columns.

Organising qualitative data in tables

For qualitative data, the column labels should include a word or phrase description of the variables. Data presented in these tables often shows trends that may be compared and contrasted in the discussion section.

Organising quantitative data in tables

For quantitative data, the column labels should include a description of the independent and dependent variables and unit(s) if measurement is involved.

Data presented in these tables displays the numerical values of each variable without a clear indication of their relationship(s). This means you need to use a graph (or other visual aid) to show any patterns or relationships.

TABLE 1 A template for organising quantitative data in a table

Independent variable (include units)	Dependent variable (include units)			
	Trial 1	Trial 2	Trial 3	Average

Graphing data

Scientific graphs have the following features in common:

- a title, which should be a descriptive statement that includes the convention 'dependent variable versus independent variable'
- axes labels with units and scale numbers
- independent variable plotted on the horizontal (x) axis
- dependent variable is plotted on the vertical (y) axis
- properly scaled axes (of equal unit size) that fit the allocated space, and which ensure that the data is plotted on the graph (not beyond each axis). The graph should be as large as possible in the allocated space for clarity
- a colour coding system or different symbols for each different set of data points if more than one data set is plotted on the same pair of axes.

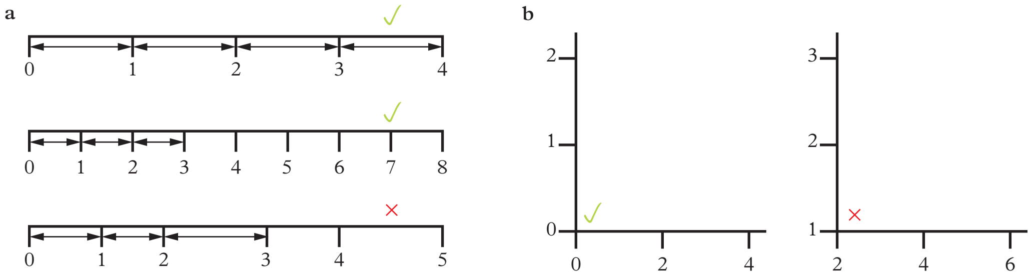


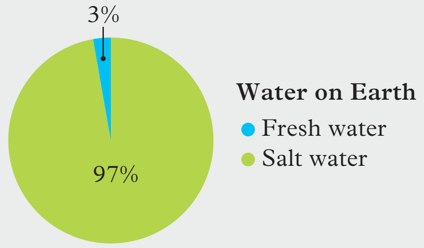
FIGURE 1 When setting up a graph, remember to **a** scale your axes evenly, and **b** begin your axes at 0 (or the origin).

Figure 1 shows some examples of what to do and what not to do when drawing graphs. The points on your axis scale should be the same distance apart. If you use numbers, they should also increase by the same amount each time.

Graphing discrete data

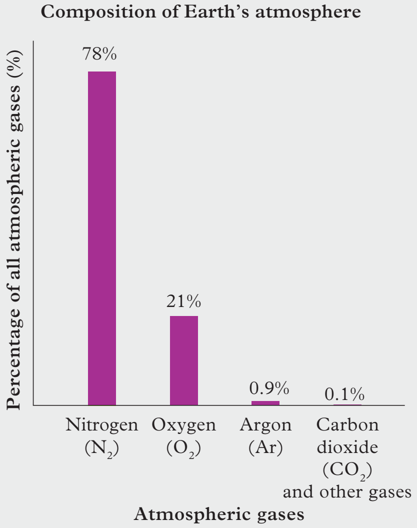
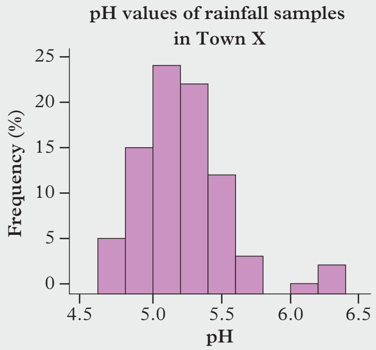
Qualitative or discrete data can be presented using different types of graphs (see Table 2).

TABLE 2 Graphing discrete data

Type of graph	Description	Example
Pie chart	Pie charts are drawn as circles and divided into segments that each represent a category. They show the proportion of each category as part of the whole data set.	

(continued)

TABLE 2 continued

Type of graph	Description	Example
Column graph	Column graphs are used to display categorical data. The bars can be plotted vertically or horizontally but should not be touching.	
Histogram	Histograms are used to show the **frequency** of numerical data groups. Each column width on a histogram can be a single data value or a data interval if the data is grouped (e.g. 0–5, 5–10, 10–15). The height of the column corresponds to the frequency, expressed as a number or percentage.	

frequency
the number of times an observation occurred in an experiment or a study

Graphing continuous data

Line graphs or scatterplots are useful ways of presenting continuous data sets where there are relationships between the dependent and independent variables.

A line graph is used to display one or more data sets that are in a defined relationship (Figure 2).

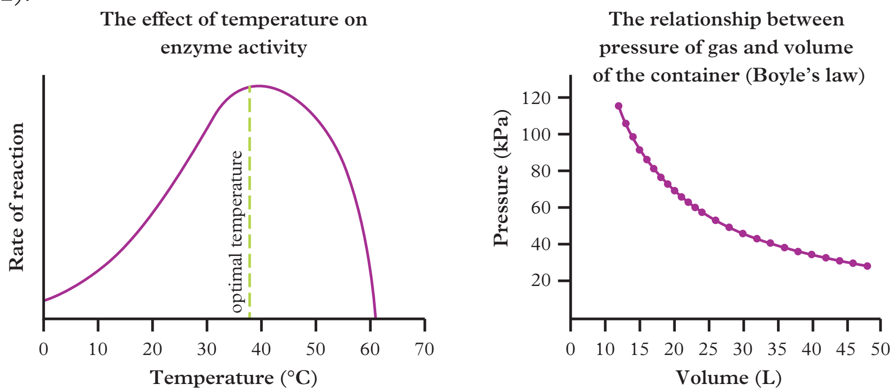


FIGURE 2 Line graphs are used to describe the relationship between two variables in a defined relationship.

A scatterplot is used to identify trends between two data sets that are not in a defined relationship.

Study tip

Do not confuse histograms with column graphs. Histograms are used for displaying numerical data, and their bars touch. Column graphs are used to show categories of data, and their bars do not touch.

Line of best fit

line of best fit
a line drawn through a scatterplot of data points that expresses a relationship between those points

A **line of best fit** can be drawn on a scatterplot to demonstrate any trends in the data (see Figure 3). Some points may be above, below or on the line of best fit. If the trend of the data is linear, the line of best fit is drawn with a straight edge. If the trend of the data is curved, a smooth curve should be drawn. If there is no visible trend between the data points (i.e. they are scattered all over the graph), then no line of best fit should be drawn.

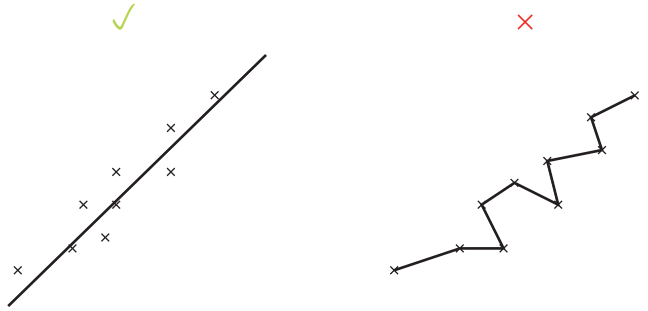


FIGURE 3 A line of best fit is drawn through the middle of your data points. You shouldn't connect the dots. A line of best fit can only be drawn if there is a strong positive or strong negative correlation.

Correlation

correlation
the presence of a relationship between two variables; can differ in strength

The relationship between two variables is called the **correlation**. If the data points on a scatterplot are close together, it could mean there is a stronger correlation between the variables.

- If one variable increases as the other variable increases, it is called a positive correlation. Weight and height are two variables that commonly have a strong positive correlation.
- If one variable increases as the other variable decreases, it is called a negative correlation.
- If there is no trend between the variables (data points are scattered all over the graph), the data is showing no correlation between the variables.

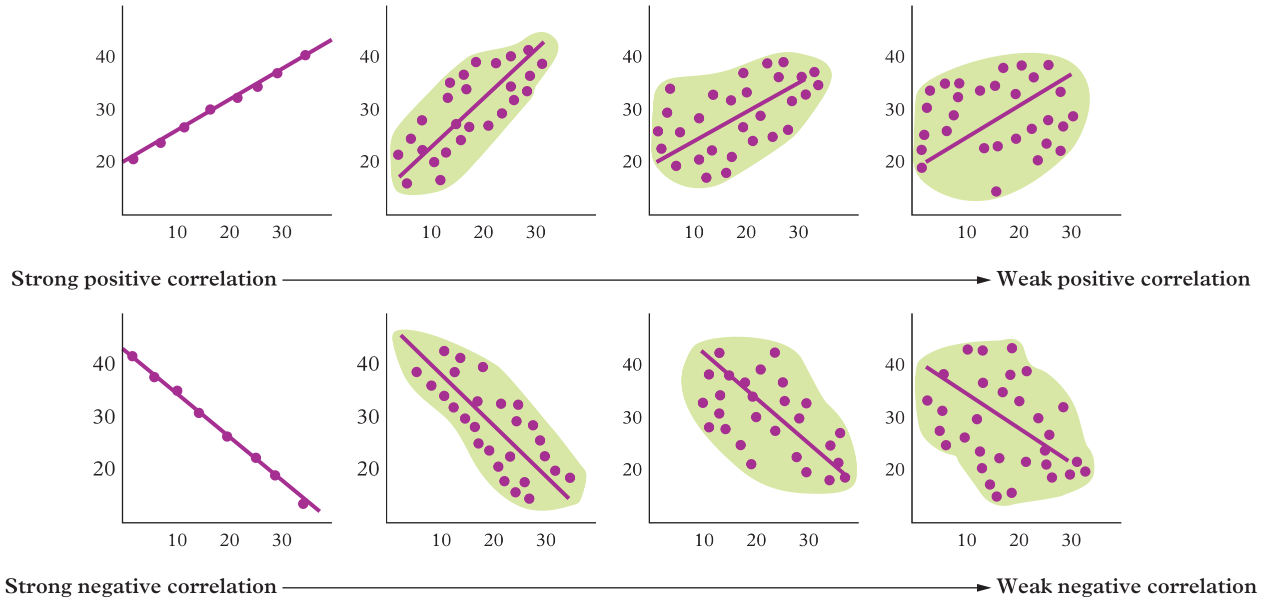


FIGURE 4 Scatterplots are used to describe the relationships between two variables that are not in a straight line.

1.7 CHECK YOUR LEARNING

Describe and explain

1 State the main features of scientific graphs.

Apply, analyse and compare

2 Contrast column graphs and histograms.

Design and discuss

3 A student conducted a redox practical investigation that simulated the measurement of blood glucose levels. They used a colorimeter and solutions of glucose ($C_6H_{12}O_6$), potassium permanganate ($KMnO_4$), distilled water and a sulfuric acid catalyst (H_2SO_4).

The student prepared a series of glucose solutions of known concentration and reacted them with a set volume of a standard solution of $KMnO_4$ (and H_2SO_4 catalyst) in 3 mL cuvettes. They also fill one cuvette with distilled water. As the permanganate ions (MnO_4^-) are reduced to manganese ions, the colour of the solutions changes from pink to colourless. Each filled cuvette is placed in the colorimeter and the students measure the percentage absorbance of each coloured solution.

FIGURE 5 A colorimeter measures the amount of light absorbed by different solutions.

The raw data obtained from this practical investigation is given in Table 3.

TABLE 3 The students' raw data

Trial	Glucose concentration (%)	Measured absorbance (%)
1	0	0.00
2	2	0.11
3	4	0.25
4	6	0.37
5	8	0.52
6	10	0.69
7	12	0.88

a Describe the purpose of using distilled water in this practical investigation.

b Identify the dependent and the independent variables in this practical investigation.

c Design a visual representation of the data in Table 3 with appropriate labels. Use the evidence from your representation and Table 3 to identify the relationship between the dependent and independent variables in this practical investigation.

1.8

Evaluating data and investigations

KEY IDEAS

In this topic, you will learn that:

- investigations that generate raw data must be valid, repeatable and reproducible
- errors and outliers must be included and accounted for in data evaluations
- significant figures must be considered in all calculations in VCE Chemistry.

The raw data generated from the design and implementation of scientific investigations must be analysed and discussed in terms of its quality and quantity; it must be **valid**, **repeatable** and **reproducible**. The VCE Chemistry Study Design provides specific guidance for terms and definitions for data, measurement, errors, significant figures and outliers.

Data and measurement

When analysing and evaluating the measurement of data, you should know a number of key terms and be able to use them. These terms are shown in bold.

When data is measured and recorded perfectly, a **true value** is obtained. If the value obtained from your experiment is close to the true value, then it is considered to be **accurate**. When the experiment is repeated and a set of replicate values is generated, you can assess the **precision** of the data by comparing how close (or similar) the values are to each other. Precision should not be confused with accuracy. It is important to understand the difference between these two concepts and be able to evaluate them. You can visualise this in Figure 1.

true value
the value you would get if the data could be measured and recorded perfectly

accuracy
how close the value is to the true value

precision
how close the data values in a set are to each other

High accuracy
High precision

Low accuracy
High precision

High accuracy
Low precision

Low accuracy
Low precision

FIGURE 1 Accuracy and precision are different concepts.

Once you determine that your data set is accurate and precise, you report this as a **measurement result**. This is usually the average of at least three trials but can also refer to a single result if only one trial was conducted.

It is also important to assess the **repeatability** of your experiment. This is whether the same data values can be produced again by the same experiment, under the same laboratory conditions. You can usually determine repeatability by conducting the same experiment at least three times.

Reproducibility is similar to repeatability. This is the ability to generate the same data values under slightly different conditions (such as with a different experiment, different measuring instrument or in a different laboratory). A reproducible experiment relies on a clear experimental method and well-defined variables. This is closely linked with accuracy.

measurement result
the final result reported in an investigation

repeatability
whether the same data values can be produced again by the same experiment, under the same laboratory conditions

reproducibility
whether the same data values can be reproduced under slightly different conditions

FIGURE 2 In these two measurement tools, the resolution of the thermometer is 1°C, and the resolution of the beaker is 25 mL.

Repeatability and reproducibility can be used to evaluate the quality or precision of measurement results.

Because different tools are used in experiments, **resolution** can be different. This is the smallest increment a tool can measure. For example, if the resolution of a burette is 0.5 mL, the experimenter needs to estimate the value if it lies between the marked lines. Measurement readings of 10.50 mL or 10.55 mL are possible, but you cannot claim a measurement reading of 10.53 mL.

Finally, you need to be able to assess the **validity** of an experiment. An experiment is considered valid if it investigates what it aims or claims to investigate. For example, if the controlled variables change or there is observer bias, the experiment would not be valid. Both experimental design and implementation should be considered when evaluating validity.

resolution
the smallest increment a tool can measure

validity
whether the experiment investigates what it aims or claims to investigate

Measurement errors, uncertainty, significant figures and outliers

Although we may aim to measure true values when conducting an investigation, this is often not achievable because of variations and **errors**. The term 'measurement error' describes the difference between a measurement result and the true value.

Conducting multiple trials of an investigation is a useful way to minimise the impact of measurement errors. By conducting more trials, you can record more data and therefore find average values for measurements across trials. The more trials you conduct, the more likely it is that the average value will be reliable.

The following list outlines different kinds of errors that might occur during your investigations.

- **Random errors** are unpredictable and affect the precision of a measurement. They are due to an error in the measurement process. For example, parallax error occurs when a measurement is different from the true value because it has been observed at the wrong angle. Random errors can be reduced by doing multiple trials and taking repeated measurements.

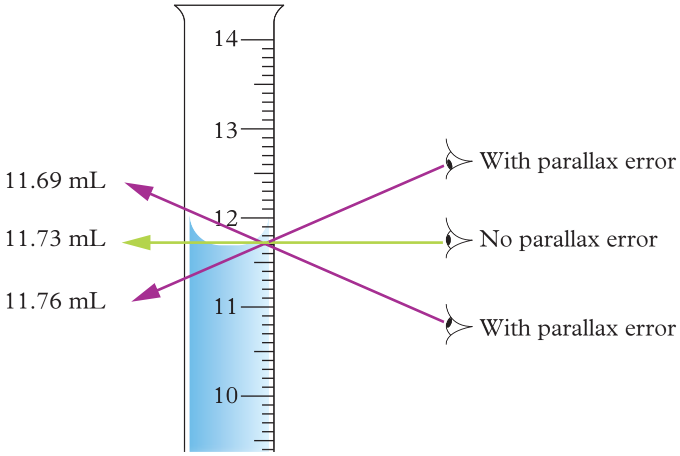


FIGURE 3 Parallax error occurs when a measurement is observed incorrectly from the wrong angle.

error
the difference between an accepted or theoretical value and the experimental, observed or measured value

Study tip

Errors are not mistakes; they are slight changes in measurement that cannot be avoided when standard operating and/or laboratory procedures are used. Common types of errors include **reading errors**, **parallax errors** or observational errors and **instrument errors**.

reading error
an error made by approximating a measurement if it falls between the markings of a measuring device

parallax error
an error made by not having your eye directly in line with the measurement

instrument error
an error due to the limited accuracy of a scientific instrument; usually, very small (and not significant) if used with standard operating and/or laboratory procedures

- **Systematic errors** are consistent and repeatable. They are caused by a problem with the method or equipment, such as an uncalibrated measuring instrument. Therefore, repeating the experiment will not reduce the impact of systematic errors.
- **Mistakes**, also called personal errors, are different from systematic or random errors. Mistakes should not be included in reporting and analysis. Instead, if you make a mistake, you need to repeat the experiment correctly. Careless measurements, or recording incorrect measurements, material spills and using incorrect equipment are examples of mistakes. Mistakes can be avoided by taking more care.
- **Uncertainty** occurs when you are unsure of the exact quantity being measured. For example, if you weigh a sample at 0.42 g, you may be uncertain if that number is 0.420 exactly, or 0.422. Each piece of equipment has a level of uncertainty.
- **Outliers** are data points or observations that are very different from the rest of the data set. Outliers need to be explained when analysing data in the discussion of a scientific report. For example, it could have been caused by an error during the investigation. Outliers should also be plotted on a graph (but they can be excluded when drawing a line of best fit). Repeating measurements can be a useful way to examine outliers.

It is important to be able to critically analyse the data that you have collected or has been provided to you. This will help you identify any contradictory or incomplete data, and recognise issues such as personal bias. This is called **data evaluation**.

data evaluation
the critical analysis of data that has been personally collected or provided to identify contradictory or incomplete data, or issues such as a personal bias

Percentage error

Percentage error refers to the difference between a measured value (e.g. one you have taken in your experiment) and a known or theoretical value. Percentage error can be a positive or negative number and is calculated as follows:

$$\%\ \text{error} = \frac{(\text{observed value}) - (\text{accepted or theoretical value})}{(\text{accepted or theoretical value})} \times 100$$

percentage error
the calculated percentage difference between an observed or measured value and an accepted or theoretical value

Significant figures

Significant figures are essential to calculations in VCE Chemistry. They are the digits of a number that are necessary to express a number accurately.

The VCAA provide the following guidelines for significant figures:

- all digits in numbers expressed in standard form are significant, e.g. 4.320×10^{-6} has 4 significant figures
- all non-zero numbers are significant, e.g. 42.3 has 3 significant figures
- zeros between two non-zero numbers are significant, e.g. 4.302 has 4 significant figures
- leading zeros are not significant, e.g. 0.0043 has 2 significant figures
- trailing zeros to the right of a decimal point are significant, e.g. 42.00 has 4 significant figures
- for numbers less than 1, 0.4 has 1 significant figure, 0.04 also has 1 significant figure whereas 0.40 has 2 significant figures and 0.400 has 3 significant figures
- whole numbers written without a decimal point will have the same number of significant figures as the number of digits, with the assumption that the decimal point occurs at the end of the number, for example 400 has 3 significant figures. Therefore, a stated volume of '400 mL' will be considered as having 3 significant figures.

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significant figures
the digits of a number that are used to express it to the required degree of accuracy, starting from the first non-zero digit

Study tip

If you are using values from the data book, remember to consider the number of significant figures in these values when you are rounding your final answer.

1.8 WORKED EXAMPLE

CALCULATING PERCENTAGE ERROR AND APPLYING SIGNIFICANT FIGURES

An experiment was performed to determine the density of lead and yields a value of 10.95 $g\,cm^{-3}$. The accepted value for the density of lead is 11.342 $g\,cm^{-3}$.

a Calculate the percentage error, showing all working.

b Round the answer to part **a** to 4 significant figures.

Solution

Think	Do
Step 1: Calculate the percentage error by substituting values into the formula: $\%\text{ error} = \frac{(\text{observed value}) - (\text{accepted or theoretical value})}{(\text{accepted or theoretical value})} \times 100$	**a** $\%\text{ error} = \frac{10.95 - 11.342}{11.342} \times 100$ $= -0.0345618 \times 100$ $= -3.45618$
Step 2: Round –3.456 18 to 4 significant figures, using VCE guidelines.	**b** –3.456%

1.8 CHECK YOUR LEARNING

Describe and explain

1 Identify the types of errors that would affect the accuracy of data collected in a scientific investigation.

2 Explain why implementing multiple trials in scientific investigations improves the reliability of data obtained.

Apply, analyse and compare

3 Compare:

- **a** accuracy and precision
- **b** repeatability and reproducibility
- **c** mistakes and errors.

4 The accepted boiling point of water is 100°C. The following boiling point values were measured by a student during an experiment under standard laboratory conditions. Comment on the accuracy and/or precision of the data obtained.

89°C, 90°C, 91°C, 88°C, 91°C, 95°C, 97°C, 93°C, 91°C, 90°C

5 Round the following values to 3 significant figures.

- **a** 4.078
- **b** 36
- **c** 02.90

Design and discuss

6 Design a revision summary tool to help you remember the application of significant figures in VCE calculations.

1.9

Constructing evidence-based arguments and conclusions

KEY IDEAS

In this topic, you will learn that:

- analysis of raw data generated in a scientific investigation is important for constructing evidence-based arguments and conclusions
- sentence structure and paragraph structure are important when constructing evidence-based arguments and conclusions.

Constructing evidence-based arguments and conclusions is an important key science skill required when studying VCE Chemistry. It is primarily used in the analysis, discussion and conclusion sections of the scientific investigation report.

How to construct evidence-based arguments and conclusions

Some broad questions that may be answered when constructing evidence-based arguments are:

- What is the origin of or who was responsible for the experimental results, findings, evidence or raw data collected?
- Are the experimental results, findings, evidence or raw data accurate, precise, reproducible and/or repeatable?
- What kinds of errors, inconsistencies and outliers may have affected the experimental results, findings, evidence or raw data collected?
- What background information (e.g. chemical concepts, scientific understandings, other researched information sources) was used to link or connect to the experimental results, learnings, evidence or raw data to the investigation question and to the aim?
- What series of steps or procedures could be used to improve the experimental design or methodology for future trials of this scientific investigation?

Structure of an argument

Argument structure usually follows a pattern, the length of which is determined by the number of arguments addressed. One argument usually corresponds to one paragraph.

Introduction (What is your argument/point of view?):

- Position statement (What is your hypothesis or summary of the scientific investigation?)
- List the arguments that you will make

Explain your supporting arguments:

- Supporting argument 1
  - Point – identified in topic sentence
  - Explanation – support with evidence, scientific finding, result or data
- Supporting argument 2:
  - Point – identified in topic sentence
  - Explanation – support with evidence, scientific finding, result or data

- Supporting argument 3:
  - Point – identified in topic sentence
  - Explanation – support with evidence, scientific finding, result or data

Reinforce your main point – ‘To sum up...’; ‘To conclude…’

An example is shown below and in Worked example 1.9.

The hypothesis of whether increasing current in an electrolysis reaction positively affected the reaction rate of electroplating, using a nickel anode and a copper cathode, was supported.

Introduction

From the line graph of the results obtained in Table 1, the mass of the nickel deposited on the copper cathode increased as the applied current increased, showing a positive linear relationship between the dependent variable (mass of nickel deposited) and independent variable (current applied). The data obtained in the experiment conformed to one of the quantitative Faraday’s Laws of electrolysis[1], in that the amount of chemical change at an electrode–electrolyte boundary (mass deposited) is directly proportional to the amount of electricity (current) used.

Argument 1, with supporting fact from external source[1]

Limitations for this experiment included that a small number of trials (one only) were conducted for each current value and the experiment design was small in scale, as the quantities recorded for the dependent and independent variables were small (all values recorded and reported were less than 1.0).

Argument 2, featuring limitations of the data collected

Suggestions for future improvement include conducting multiple trials for each current setting to obtain more reliable values for the mass of nickel deposited and to build a larger-scale electrolytic cell to take larger current input.

Argument 3, featuring suggestions for improvement

In conclusion, increasing the amount of current in an electrolysis reaction increased the reaction rate of electroplating of a nickel anode onto a copper cathode in the above experiment. Conducting repeated trials on a larger-sized electrolytic cell would improve upon the reliability of the data obtained.

Conclusion

1 Britannica, The Editors of Encyclopaedia (March 2021) ‘Faraday’s laws of electrolysis’ Accessed 25th January 2022 from https://www.britannica.com/science/Faradays-laws-of-electrolysis

Language features

When writing an argument, use the following language features to enhance your writing:

- Use connectives to show cause and effect; for example, ‘As an outcome of ...’, ‘As a result of ...’, ‘because ...’, ‘consequently ...’
- Use scientific terminology and/or supporting figures.
- Use supporting facts and/or quotes from experts or other researched external sources of information; for example, ‘A study conducted by XX found that …’
- Make evaluative statements; for example, ‘The data shows a clear trend …’, ‘Evidence contradicts the argument that …’
- Use formal language and avoid personal language (such as I, you, he, she, we, they, me, him, her, us and them); for example, ‘the hypothesis was supported’, not ‘I proved my hypothesis’.

1.9 Worked example
Find me in your obook pro

1.9 Worked example
Video demonstration

1.9 CHECK YOUR LEARNING

Describe and explain

1 Explain why it is important to construct and understand evidence-based arguments and conclusions in science.

2 Identify three things you should (or should not) do when writing an evidence-based argument.

Design and discuss

3 Design an evidence-based analysis and conclusion for the experimental data presented in the ‘Check your learning’ Topic 1.7, Question 3.

1.10 Communicating

KEY IDEAS

In this topic, you will learn that:

+ effective science communication depends on knowing your audience and prioritising important information.

Science communication is not just found in your science subjects at school. It is found in everyday life: online (websites, blogs, videos, social media), in journalism, in advertising and in government policy.

Important points to remember in order to communicate science effectively:

- **Identify and understand your target audience** – presenting work for teachers or external examiners will be different from presenting to your peers.
- **Use language that is appropriate for your target audience** – if you are speaking to the general public, then limit the use of **jargon** wherever possible. If you need to incorporate **acronyms**, make sure that the full words are written in brackets next to the acronym on one occasion.
- **Only include essential content** – your communication should contain essential facts that enhance your target audiences' understanding of the science involved.
- **Use appropriate stylistic elements that are relevant to your target audience** – this may include (but is not limited to) the use of a specific presentation format (poster, report, infographic or digital media platform), diagrams, photographs, graphs, tables or other text elements (e.g. similes, metaphors or analogies). They should convey clear content and messages, and they should appeal to your audience and enhance the information so that the science is simple, easily understood and relatable without excess information or confusion.
- **Use appropriate scientific literacy and conventions** – this may include (but is not limited to) the use of coherent, concise, objective and formal use of scientific language; the consistent use of tense and/or active or passive voice; definition of key scientific terms and accurate scientific representations (including models, symbols, units, balanced chemical equations and/or formulas) and a bibliography or reference list for citation of text and/or diagrams that are not of your own creation.

jargon
special words or expressions used by a profession or group that are difficult for others to understand

acronym
an abbreviation formed from the initial letters of each word from a group of words

Study tip

Spell check on your digital device is not always 100% reliable. Read, re-read or ask another person to review your work before final submission(s) so that the points of effective science communication have been covered.

Formats for communication

Science is not all about practical reports. There are many ways to communicate scientific concepts that you might like to try in VCE Chemistry. For example, you could present information as:

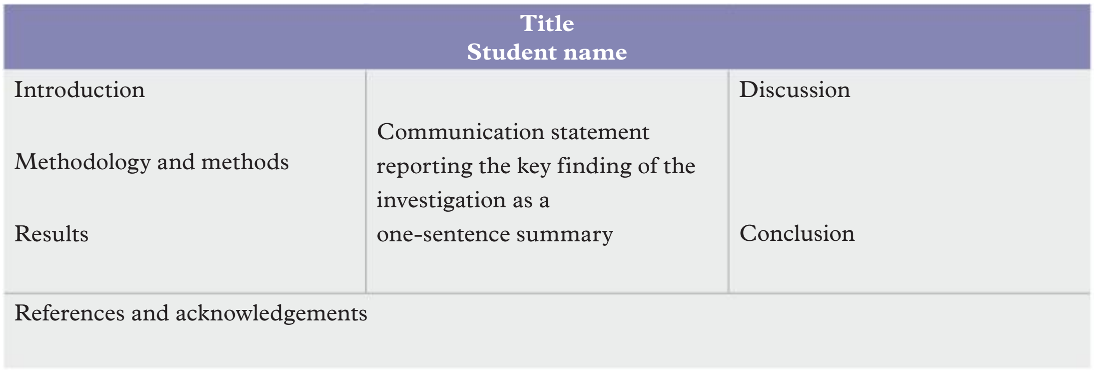
- an oral presentation
- a report
- an infographic
- a video or animation
- a multimodal presentation
- a scientific poster.

Scientific poster

In VCE Chemistry, you will produce a scientific poster as part of a major piece of assessment in Units 3 and 4.

The poster may be produced electronically or in hard copy and should not exceed 600 words, to meet criteria of conciseness, clarity and legibility. Note that tables, graphs and all references and/or acknowledgments are not included in the word count.

Figure 1 shows the required scientific poster format as specified by the VCAA Study Design. The centre of the poster should take up 20–25% of the poster space and contain a one-sentence summary of the major finding or outcome of the investigation that answers the investigation question.



Title
Student name

Introduction

Methodology and methods

Results

Communication statement reporting the key finding of the investigation as a one-sentence summary

Discussion

Conclusion

References and acknowledgements

FIGURE 1 The format for the scientific poster

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1.10 CHECK YOUR LEARNING

Describe and explain

1 **a** Define some important elements of effective science communication.

b Describe which of these elements would be most effective for presenting your school-assessed coursework.

2 Explain why it is important to use appropriate scientific language and literacy throughout the poster.

Apply, analyse and compare

3 Compare the elements of science communication used in a science video and a media article of choice. Describe the ways in which they are similar and different.

Design and discuss

4 Compare the results and the discussion sections of the poster.

5 Using appropriate science communication elements, design a presentation (of choice) for the experimental data, evidence-based analysis and conclusion created in the ‘Check your learning’ Topics 1.7 and 1.9, Question 3.

1.11

Sustainability

KEY IDEAS

In this topic, you will learn that:

+ studies of sustainability have been integrated into Units 1–4 of the VCE Chemistry Study Design.

sustainability
using natural resources more efficiently, to ensure supply of resources is continued into the future

Sustainability in chemistry refers to the efficiency with which we use natural resources to meet society's needs for chemical products and services. Sustainability is important so that we do not use up all of Earth's resources, leaving nothing for future generations.

It is a key focus of the VCE Chemistry Study Design. Where possible in VCE Chemistry, you should consider how you can link ideas to sustainability.

You will consider sustainability in terms of three main perspectives:

- sustainable development
- green chemistry principles
- linear and circular economies.

In this topic, you will explore the three sustainability perspectives and see how and where they fit into VCE Chemistry. These perspectives are explored throughout the student book. This includes school-based assessments for Areas of Study 1 or 2, where you will need to reference sustainability in at least one of them.

Study tip

Further detail on the United Nations Sustainable Development Goals is found on the United Nations website (link via your obook pro).

Sustainable development

The term *sustainable development* is defined in a report by the World Commission on Environment and Development as development that meets the needs of people in the present without affecting the ability of people in the future to meet their needs.

This led the United Nations to identify their own Development Goals (see Table 1 and Figure 1) that address current global challenges for sustainability.

For VCE Chemistry, you need to be aware of nine of these goals and how they can be applied.

FIGURE 1 The 17 United Nations Sustainable Development Goals

TABLE 1 The nine United Nations Sustainable Development Goals relevant to VCE Chemistry and where they fit into Units 3 and 4

Goal	What would it look like to achieve this goal?	Student book chapter
2 Zero hunger	End hunger, achieve food security, improve nutrition and promote sustainable agriculture	2 (Topic 2.4), 9 (Topics 9.2, 9.4), 13
6 Clean water and sanitation	Provide sustainable access to water and sanitation to everyone	10, 13
7 Affordable and clean energy	Provide access to affordable, reliable, sustainable and modern energy for everyone	2 (Topics 2.1, 2.5), 3, 4 (Topic 4.7), 7 (Topic 7.4), 13, 14 (Literature review 2.1, Literature review 3.1, Literature review 4.7)
9 Industry, innovation and infrastructure	Build resilient infrastructure, use inclusive and sustainable methods of building and production, foster innovation (i.e. new ideas)	2 (Topic 2.5), 4 (Topic 4.7), 5, 6 (Topics 6.4, 6.7), 7 (Topics 7.4, 7.5), 9 (Topics 9.3, 9.4), 13, 14 (Literature review 2.1, Literature review 3.1, Literature review 4.7, Controlled experiment 5.1, Simulation 5.2, Case study 6.3B, Simulation 6.7A, Product, process or system development 6.7B, Controlled experiment 7.1)
11 Sustainable cities and communities	Make cities and human settlements inclusive, safe, resilient and sustainable	2 (Topic 2.5), 4 (Topic 4.7), 7 (Topic 7.4), 9 (Topic 9.4), 14 (Literature review 2.1, Literature review 3.1, Literature review 4.7)
12 Responsible consumption and production	Ensure resources are consumed and produced sustainably	2 (Topics 2.1, 2.5, 2.8), 3 (Topics 3.1, 3.4), 4 (Topic 4.7), 6 (Topics 6.4, 6.7), 7 (Topic 7.4), 9 (Topics 9.3, 9.4), 13, 14 (Literature review 2.1, Literature review 3.1, Literature review 4.7, Controlled experiment 5.1, Simulation 5.2, Case study 6.3B)
13 Climate action	Take urgent action to combat the climate crisis, including drastically reducing greenhouse gas emissions	2 (Topics 2.1, 2.5), 3, 4 (Topic 4.7), 7 (Topic 7.4), 13, 14 (Literature review 2.1, Literature review 3.1, Literature review 4.7, Case study 6.3B)
14 Life below water	Protect and use sustainably the oceans, seas and marine resources	13, 14 (Case study 6.3B)
15 Life on land	Protect, restore and promote ecosystems on land, manage forests, combat desertification, and reverse land degradation and loss of biodiversity	2 (Topic 2.5), 13

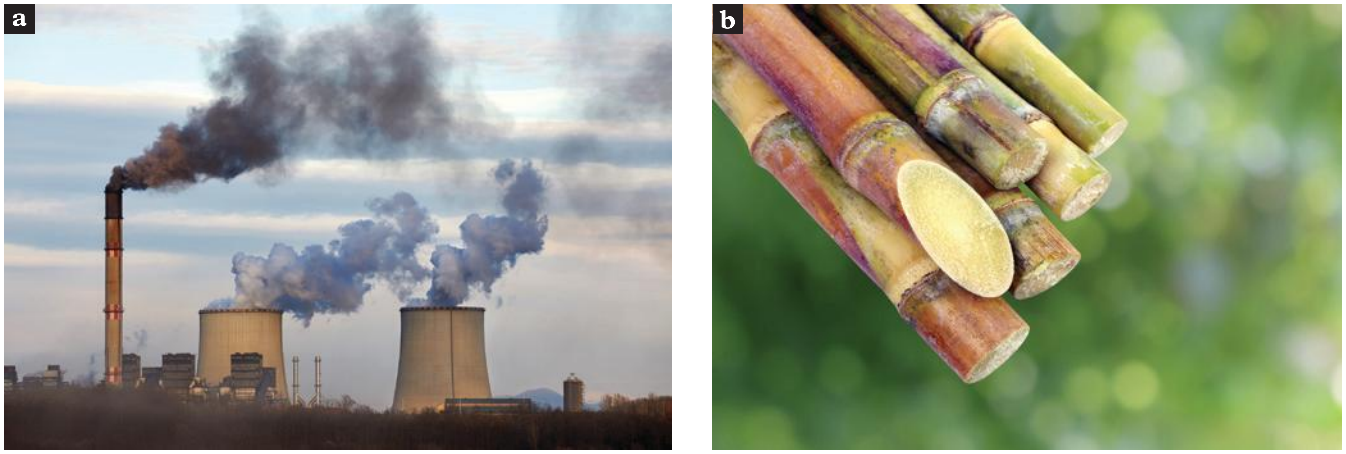


FIGURE 2 United Nations Sustainable Development Goal 12 is a big focus in Units 3 and 4. You will learn a lot about shifting away from **a** fossil fuels and towards **b** fuels derived from biomass, like cane sugar.

Green chemistry principles

green chemistry
an area of chemistry that focuses on designing safer and more sustainable new products

Green chemistry is an area of chemistry that focuses on designing safer and more sustainable new products. It is based on several principles that aim to reduce the impact of product or process development on the environment, resources and human well-being.

The principles focus on reducing risks or hazards, the amount of materials consumed, the amount of energy used and unwanted wastes.

To achieve these aims, green chemistry relies on creativity and innovation in developing new products, processes and systems.

Seven green chemistry principles are important in VCE Chemistry (Table 2).

TABLE 2 The seven green chemistry principles important in VCE Chemistry and where they fit into Units 3 and 4

Green chemistry principle	Description	Student book chapter
Atom economy	Atom economy refers to how much of the initial reactant atoms are contained in the final product. Processes that are designed to have good atom economy maximise the incorporation of reactant materials into the final product.	2 (Topic 2.3), 4 (Topic 4.3), 10, 19 (Literature review 2.3, Fieldwork 4.3B)
Catalysis	Catalysis involves using a catalyst (such as an enzyme) to increase the rate of a chemical reaction. In green chemistry, the use of catalysis aims to generate the same product with less energy or waste in a reaction.	10
Design for degradation	Under this principle, products are designed to break down (degrade) into harmless products at the end of their life cycle. They do not persist in or harm the environment.	9, 13
Design for energy efficiency	Under this principle, processes are designed to increase energy efficiency.	5, 6 (Topic 6.7), 7 (Topic 7.4), 13
Designing safer chemicals	Under this principle, chemical products are designed to be less harmful or toxic, yet still able to carry out their purpose.	9 (Topic 9.4), 13
Prevention of wastes	This principle seeks to prevent waste rather than treating or cleaning it up after it has been created.	7, 9 (Topic 9.3), 13, 14 (Literature review 3.1)
Use of renewable feedstocks	A raw material or a **feedstock** (used for processing another product) should be made from sustainable materials (e.g. plant-based), rather than fossil fuels.	2 (Topics 2.1, 2.5), 4 (Topic 4.7), 9 (Topic 9.4), 12 (Topic 12.1), 13, 14 (Literature review 2.1, Literature review 3.1, Literature review 4.7)

feedstock
a raw material used to produce other goods

linear economy
a way of managing resources that operates on a take–make–dispose model; new resources are used and disposed of after use

Linear and circular economies

Natural resources are becoming scarce. Humans are using and disposing of more products than we have ever done in the past. This is due to growing global populations, but also due to global spending habits.

As a society, we must rethink the way we consume and dispose of products to make sure our natural resources are available in the future.

In VCE Chemistry, you may be asked to compare how a **linear economy** and a **circular economy** make use of our resources.

Linear economy

A linear economy uses a ‘take–make–dispose’ model. Resources are extracted from nature or the Earth to make products that will be thrown away (sometimes after only one use). A simple representation is shown in Figure 3.

circular economy
a model of production that involves sharing, leasing, re-using, repairing, refurbishing and recycling existing materials and products as long as possible

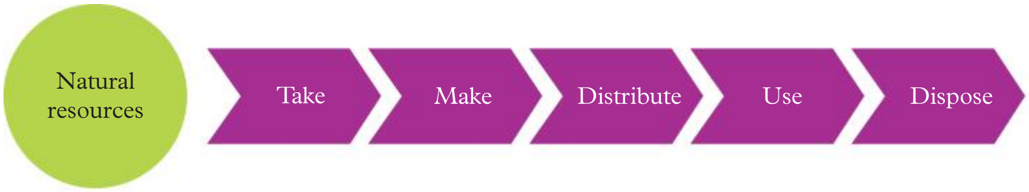


FIGURE 3 A representation of a linear economy

Circular economy

A circular economy is a more sustainable approach. It follows a continuous cycle that focuses on using and re-using resources repeatedly at different stages of production and consumption (Figure 4).

Linear and circular economy concepts are explored in the following chapters in the student book:

- Chapter 2 (Topic 2.5)
- Chapter 7
- Chapter 8 (Topics 8.3, 8.4)
- Chapter 9
- Chapter 10
- Chapter 13.

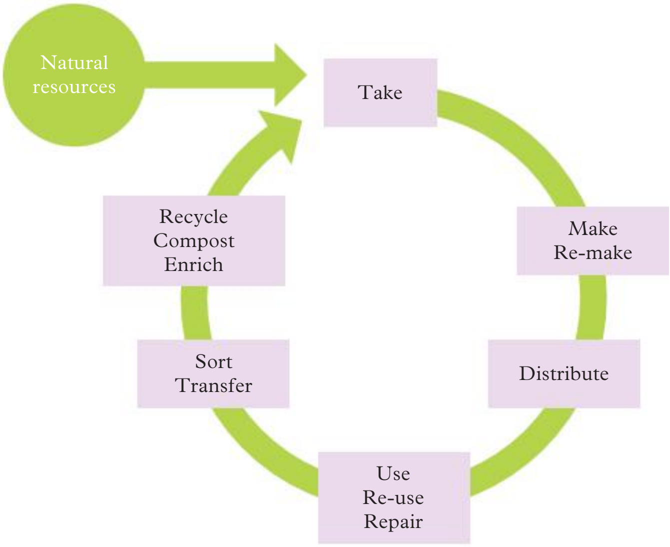


FIGURE 4 A representation of a circular economy

1.11 CHECK YOUR LEARNING

Describe and explain

1 Describe and explain which principles of green chemistry are the most relevant to studying VCE Chemistry.

2 Explain why the transition from a linear to a circular economy is being used as a strategy for more sustainable development.

Design and discuss

3 Some hazardous substances that have been minimised or eliminated by green chemistry processes include carbonyl chloride (for the manufacture of bisphenol A in plastics), sodium salts of alkyl benzene sulfonic acids (for detergents), sodium phosphate builders (for removing calcium and magnesium ions from hard water), chlorofluorocarbons (for refrigeration, air conditioning and aerosol propellants) and phosphorus- or chlorine-based pesticides (for crop pest control). Select one of the hazardous substances mentioned and research how green chemistry was used to replace it. Present your findings to your class.

1.12 Using the periodic table

KEY IDEAS

In this topic, you will learn that:

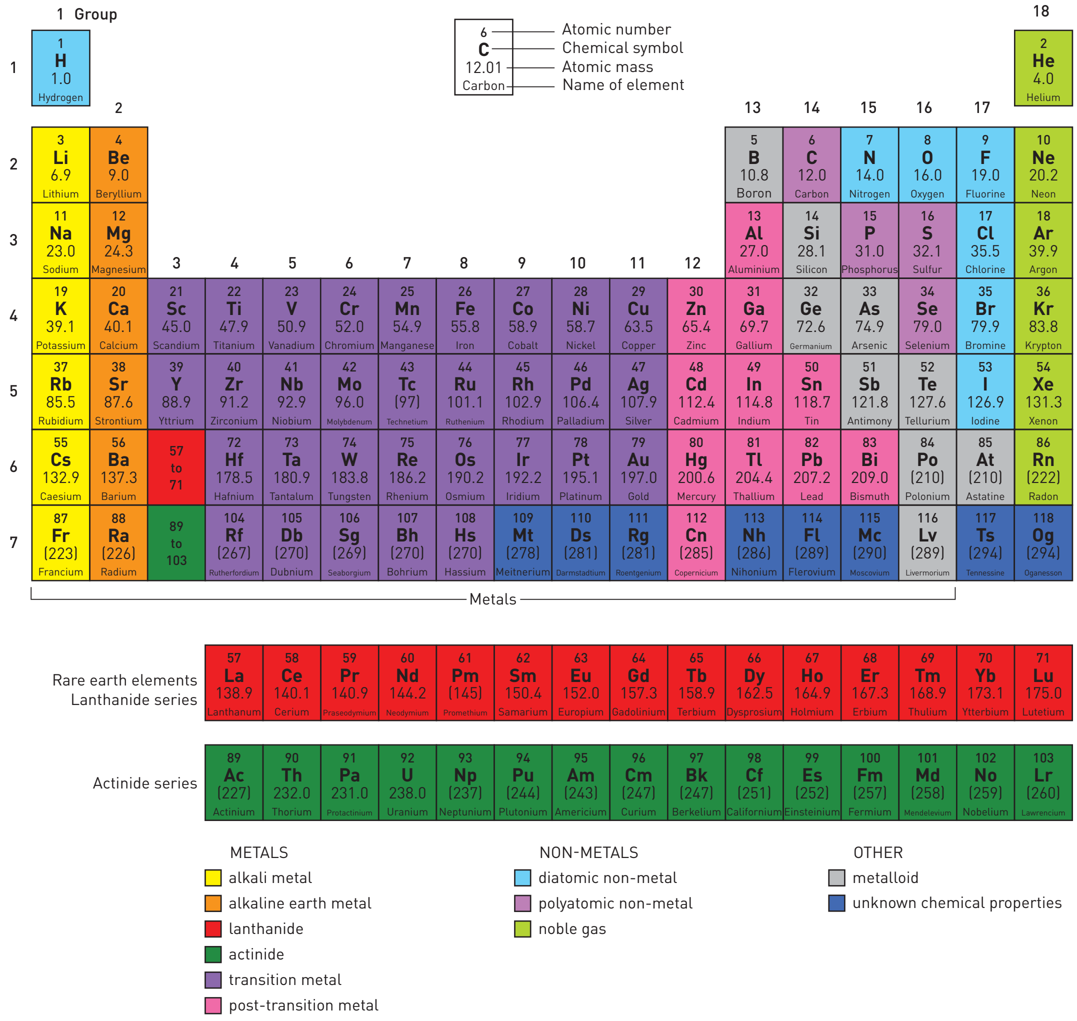
- all known elements are arranged in rows and columns in the periodic table according to their electron shell configuration and properties
- the periodic table is an important learning tool in VCE Chemistry.

periodic table
the table that organises the elements in increasing order of atomic number

proton
a subatomic particle in an atom's nucleus with a positive charge

The early development of the **periodic table** is credited to the Russian scientist Dmitri Mendeleev. In 1869, he arranged around 60 known elements based on their reactivities.

The modern periodic table (Figure 1) arranges the elements in order of (from left to right) increasing atomic number or number of **protons**. You should remember from Units 1 and 2 that elements in the periodic table are arranged in different families: alkali metals, alkaline earth metals, transition metals, post-transition metals, lanthanides, actinides, metalloids, non-metals (diatomic and polyatomic) and noble gases. These have been colour-coded and labelled for your reference.



6 — Atomic number
C — Chemical symbol
12.01 — Atomic mass
Carbon — Name of element

Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.0 Hydrogen																	2 He 4.0 Helium
2	3 Li 6.9 Lithium	4 Be 9.0 Beryllium											5 B 10.8 Boron	6 C 12.0 Carbon	7 N 14.0 Nitrogen	8 O 16.0 Oxygen	9 F 19.0 Fluorine	10 Ne 20.2 Neon
3	11 Na 23.0 Sodium	12 Mg 24.3 Magnesium											13 Al 27.0 Aluminium	14 Si 28.1 Silicon	15 P 31.0 Phosphorus	16 S 32.1 Sulfur	17 Cl 35.5 Chlorine	18 Ar 39.9 Argon
4	19 K 39.1 Potassium	20 Ca 40.1 Calcium	21 Sc 45.0 Scandium	22 Ti 47.9 Titanium	23 V 50.9 Vanadium	24 Cr 52.0 Chromium	25 Mn 54.9 Manganese	26 Fe 55.8 Iron	27 Co 58.9 Cobalt	28 Ni 58.7 Nickel	29 Cu 63.5 Copper	30 Zn 65.4 Zinc	31 Ga 69.7 Gallium	32 Ge 72.6 Germanium	33 As 74.9 Arsenic	34 Se 79.0 Selenium	35 Br 79.9 Bromine	36 Kr 83.8 Krypton
5	37 Rb 85.5 Rubidium	38 Sr 87.6 Strontium	39 Y 88.9 Yttrium	40 Zr 91.2 Zirconium	41 Nb 92.9 Niobium	42 Mo 96.0 Molybdenum	43 Tc (97) Technetium	44 Ru 101.1 Ruthenium	45 Rh 102.9 Rhodium	46 Pd 106.4 Palladium	47 Ag 107.9 Silver	48 Cd 112.4 Cadmium	49 In 114.8 Indium	50 Sn 118.7 Tin	51 Sb 121.8 Antimony	52 Te 127.6 Tellurium	53 I 126.9 Iodine	54 Xe 131.3 Xenon
6	55 Cs 132.9 Caesium	56 Ba 137.3 Barium	57 to 71	72 Hf 178.5 Hafnium	73 Ta 180.9 Tantalum	74 W 183.8 Tungsten	75 Re 186.2 Rhenium	76 Os 190.2 Osmium	77 Ir 192.2 Iridium	78 Pt 195.1 Platinum	79 Au 197.0 Gold	80 Hg 200.6 Mercury	81 Tl 204.4 Thallium	82 Pb 207.2 Lead	83 Bi 209.0 Bismuth	84 Po (210) Polonium	85 At (210) Astatine	86 Rn (222) Radon
7	87 Fr (223) Francium	88 Ra (226) Radium	89 to 103	104 Rf (267) Rutherfordium	105 Db (270) Dubnium	106 Sg (269) Seaborgium	107 Bh (270) Bohrium	108 Hs (270) Hassium	109 Mt (278) Meitnerium	110 Ds (281) Darmstadtium	111 Rg (281) Roentgenium	112 Cn (285) Copernicium	113 Nh (286) Nihonium	114 Fl (289) Flerovium	115 Mc (290) Moscovium	116 Lv (289) Livermorium	117 Ts (294) Tennessine	118 Og (294) Oganesson

Metals

Rare earth elements Lanthanide series	57 La 138.9 Lanthanum	58 Ce 140.1 Cerium	59 Pr 140.9 Praseodymium	60 Nd 144.2 Neodymium	61 Pm (145) Promethium	62 Sm 150.4 Samarium	63 Eu 152.0 Europium	64 Gd 157.3 Gadolinium	65 Tb 158.9 Terbium	66 Dy 162.5 Dysprosium	67 Ho 164.9 Holmium	68 Er 167.3 Erbium	69 Tm 168.9 Thulium	70 Yb 173.1 Ytterbium	71 Lu 175.0 Lutetium
Actinide series	89 Ac (227) Actinium	90 Th 232.0 Thorium	91 Pa 231.0 Protactinium	92 U 238.0 Uranium	93 Np (237) Neptunium	94 Pu (244) Plutonium	95 Am (243) Americium	96 Cm (247) Curium	97 Bk (247) Berkelium	98 Cf (251) Californium	99 Es (252) Einsteinium	100 Fm (257) Fermium	101 Md (258) Mendelevium	102 No (259) Nobelium	103 Lr (260) Lawrencium

METALS
🟨 alkali metal
🟧 alkaline earth metal
🟥 lanthanide
🟩 actinide
🟪 transition metal
▨ post-transition metal

NON-METALS
🟦 diatomic non-metal
▨ polyatomic non-metal
▨ noble gas

OTHER
⬜ metalloid
🟦 unknown chemical properties

FIGURE 1 The periodic table

group
a vertical column on the periodic table

electron
a subatomic particle in an atom's nucleus with a negative charge

The 18 vertical columns of the periodic table are called **groups**. Elements in the same group have very similar chemical and physical properties. They also have the same number of **electrons** in their outermost or **valence shell**.

The seven horizontal rows of the periodic table are called **periods**. The period number indicates the number of electron shells. For example, elements in the fourth period (or row) of the periodic table contain four electron shells.

valence shell
the outer shell of an atom where electrons are found

period
a horizontal row on the periodic table

Using the periodic table

Each rectangle of the periodic table contains useful information for constructing and balancing chemical equations (see Topic 1.13) and for all quantitative calculations required during studies of VCE Chemistry.

From Figure 2, you can find the atomic number or the number of protons at the top of the rectangle. Every element has a unique atomic number. This value is also equal to the number of electrons of the uncharged (or neutral) element.

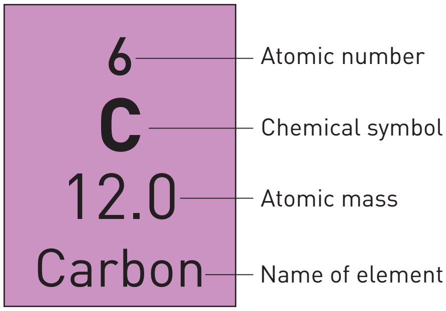


FIGURE 2 The element carbon has an atomic number of 6, atomic mass of 12.0 and chemical symbol C.

Under the atomic number, you will see the chemical symbol of the element. The convention for writing chemical symbols is either a capital letter by itself (like C for carbon in Figure 2) or a capital letter followed by a lower-case letter (like Co for cobalt, one of the transition metals).

Beneath the chemical symbol is the atomic mass of the element. This is the sum of the number of protons and **neutrons** in an element. For most elements, this is larger than the atomic number and is not a whole number (for example, 24.3 for magnesium). The atomic mass is equal to the weighted average of its individual isotopic masses, which depends on the natural **abundance** of the element. Remember from Units 1 and 2 that isotopes of an element have the same number of protons (atomic number) but different numbers of neutrons.

neutron
a subatomic particle in an atom's nucleus with a neutral (or zero) charge

abundance
how often the substance occurs

The full name of the chemical element is shown at the bottom of each rectangle, like carbon from Figure 2.

Trends of the periodic table

Patterns of physical and chemical properties can be read from the periodic table.

Some of the major trends of the periodic table going down a group (vertically) are summarised in Table 1. Many of these trends can be explained by the number of electron shells.

In Table 2 (on the next page), you can find some of the major trends on the periodic table going across a period (horizontally). Many of these trends can be explained by the **nuclear charge**.

nuclear charge
the total charge of the nucleus due to the number of positively charged protons

shielding effect
the shielding of valence electrons by inner electrons that alters the nuclear charge felt by valence electrons

TABLE 1 Periodic table trends going **down** a group (top to bottom)

Property	Trend	Reason
Atomic size or radius	increases	More electrons are present. Therefore, more space is required to accommodate them.
Metallic character (or how readily an atom loses a valence electron)	increases	Valence electrons are further away from the nucleus and can be removed more easily or can become delocalised to form an 'electron sea'.
Reducing strength of metals (or the ability to lose electrons)	increases	Valence electrons are further away from the nucleus and can be removed more easily.
Oxidising strength of non-metals (or the ability to gain electrons)	decreases	Valence electrons are further away from the nucleus, making it more difficult to attract (or gain) new electrons.

(continued)

TABLE 1 continued

Property	Trend	Reason
First ionisation energy (or the amount of energy in the gaseous state required to lose an electron)	decreases	Valence electrons are further away from the nucleus. Therefore, the energy required to remove one of these electrons is reduced.
Electronegativity (the ability of a nucleus to attract electrons towards itself)	decreases	Valence electrons are further away and are not as strongly attracted to the nucleus. This is also due to an increased **shielding effect**, where a greater number of inner electron shells reduces the charge experienced by valence electrons.
Reactivity (depending on how easily electrons are transferred to/from the atom)	increases	Depends on first ionisation energy and electronegativity. As both decrease going down a group, the atom is more likely to gain or lose electrons to form a new bond in a chemical reaction.

TABLE 2 Periodic table trends going **across** a period (left to right)

Property	Trend	Reason
Atomic size or radius	decreases	Nuclear charge increases, attracting electrons closer. This decreases the volume or size of the atom. *(This holds roughly true* ***but*** *there are some exceptions to this pattern.)*
Metallic character	decreases	Nuclear charge increases, attracting valence electrons closer. It is therefore more difficult to lose electrons to form an 'electron sea'.
Reducing strength of metals	decreases	Nuclear charge increases, attracting valence electrons closer. It is therefore more difficult to lose electrons.
Oxidising strength of non-metals	increases	Nuclear charge increases, attracting valence electrons closer. This makes it more difficult to attract (or gain) new electrons.
First ionisation energy	increases	Nuclear charge increases, attracting valence electrons closer. Therefore, more energy is required to remove a valence electron.
Electronegativity	increases	Nuclear charge increases, attracting valence electrons closer.
Reactivity	decreases	Depends on first ionisation energy and electronegativity. As both increase going across a period, the atom is less likely to gain or lose electrons to form a new bond in a chemical reaction. *(This holds roughly true* ***but*** *there are exceptions to this pattern.)*

1.12 CHECK YOUR LEARNING

Describe and explain

1 Explain the trend in electronegativity of the non-metals.

2 Explain the trend in the metallic character of the transition metals, with reference to hafnium (Hf) and copper (Cu) in your response.

Apply, analyse and compare

3 Use the periodic table to complete the missing information below.

Atom	Name of element	Number of protons	Number of neutrons
$^{32}_{15}P$			
	Uranium		
		31	32

4 Compare the positions of caesium and fluorine in the periodic table and use the trends summarised in Tables 1 and 2 to predict their relative reactivities.

1.13

Balancing chemical equations

KEY IDEAS

In this topic, you will learn that:

- balanced chemical equations obey the law of conservation of mass
- construction and balancing of chemical equations is a prerequisite for completing quantitative calculations in VCE Chemistry.

Constructing and balancing chemical equations is an important skill in VCE Chemistry. School assessment questions often ask you to 'write the balanced chemical equation for the reaction between …' before you start the quantitative calculations.

The law of conservation of mass

law of conservation of mass
a law that states that in an isolated system, mass cannot be created or destroyed

isolated system
a thermodynamic system that cannot exchange matter or energy outside its boundaries

The **law of conservation of mass** states that in an **isolated system**, during a chemical reaction, mass is conserved. It cannot be created or destroyed.

This means that in chemical reactions, the chemical bonds between the reactants and products are broken and rearranged, so that the number of atoms involved remains unchanged. Balanced chemical equations obey the law of conservation of mass. But the number of atoms present does not change.

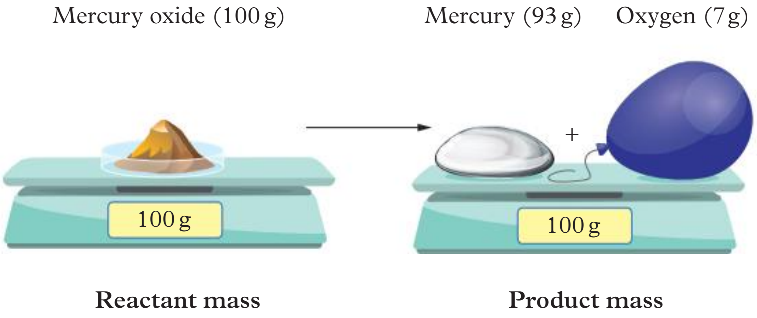


FIGURE 1 When mercury oxide (HgO) is heated, its atoms rearrange to form liquid mercury (Hg) and oxygen gas (O_2). The mass of the reactants and products is the same.

This concept is important to remember when balancing equations. There needs to be the same number of molecules on the reactant side as the product side.

Steps for balancing chemical equations

You can follow these steps when asked to balance a chemical equation.

Step 1 Write the equation with the known products and reactants

Write the chemical formulas for the reactants and products. These formulas cannot be changed to make the equation balance (i.e. do not change the subscripts).

$$CH_4 + O_2 \rightarrow H_2O + CO_2$$

This equation is **not** balanced.

Step 2 Tally the atoms

Tally how many atoms of each element are on either side of the arrow.

CH_4	+	O_2	→	H_2O	+	CO_2
C = 1		C = 0		C = 0		C = 1
H = 4		H = 0		H = 2		H = 0
O = 0		O = 2		O = 1		O = 2

C = 1 + 0 = 1	C = 0 + 1 = 1
H = 4 + 0 = 4	H = 2 + 0 = 2
O = 0 + 2 = 2	O = 1 + 2 = 3

In this example, the number of carbon atoms is balanced (one on each side) but the numbers of hydrogen and oxygen atoms are not balanced.

Step 3 Add coefficients

Put numbers (coefficients) in **front** of the molecules to add further molecules. These numbers multiply all the atoms in the chemical formula (i.e. putting a 2 in front of H_2O means there are now twice as many H and twice as many O atoms).

Balance one element at a time.

Balance hydrogen, by adding the coefficient 2 to the product side.

$$CH_4 + O_2 \rightarrow 2H_2O + CO_2$$

CH_4	+	O_2	→	H_2O	+	CO_2
C = 1 + 0 = 1				C = 0 + 1 = 1		
H = 4 + 0 = 4				H = 2 + 0 = 2 × 2 = 4		
O = 0 + 2 = 2				O = 1 × 2 + 2 = 4		

Balance oxygen by adding the coefficient 2 to the reactant side.

$$CH_4 + 2O_2 \rightarrow 2H_2O + CO_2$$

CH_4	+	O_2	→	H_2O	+	CO_2
C = 1 + 0 = 1				C = 0 + 1 = 1		
H = 4 + 0 = 4				H = 2 + 0 = 2 × 2 = 4		
O = 0 + 2 × 2 = 4				O = 1 × 2 + 2 = 4		

There are now equal amounts of C, H and O on both sides of the equation.

Step 4 Check your work

Check your work by re-counting all types of atoms on both sides of the arrow (this equation is now **balanced**).

Figure 2 is a visual representation for this chemical reaction.

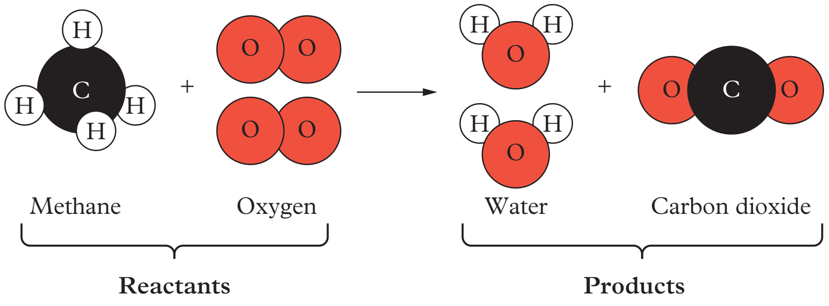


FIGURE 2 The reactants (methane and oxygen) have the same number of C, H and O atoms as the products (water and carbon dioxide) of this reaction.

Step 5 Add the states

If you know the states, place the following symbols after the formulas; (g) for a gas, (l) for a liquid, (s) for a solid, (aq) for an aqueous solution (in water).

For the previous equation for the complete combustion of methane, we would write:

$$CH_4(g) + 2O_2(g) \rightarrow 2H_2O(g) + CO_2(g)$$

In this case, the methane is being burned at a high temperature, so the water will be in a gaseous state.

1.13 CHECK YOUR LEARNING

Describe and explain

1 Explain why balanced chemical equations obey the law of conservation of mass.

Apply, analyse and compare

2 Balance the following chemical equations:

- **a** $Na(s) + Cl_2(g) \rightarrow NaCl(s)$
- **b** $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$
- **c** $Ag_2CO_3(s) \rightarrow Ag_2O(s) + CO_2(g)$
- **d** $C_2H_4(g) + O_2(g) \rightarrow CO(g) + H_2O(g)$
- **e** $P_4(s) + O_2(g) \rightarrow P_2O_5(s)$
- **f** $Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$
- **g** $HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$
- **h** $Ca(CN)_2(s) + HCl(aq) \rightarrow HCN(aq) + CaCl_2(aq)$
- **i** $Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$
- **j** $Mg(s) + O_2(g) \rightarrow MgO(s)$

Design and discuss

3 Using Figure 2 as an example, design a visual representation of one of the balanced chemical equations in Question 2.

FIGURE 3 A balanced equation has the same number of each type of atom on either side of the arrow.

1.14 Preparing for assessment

KEY IDEAS

In this topic, you will learn that:

- organisational skills are important for revision
- understanding the meaning of command terms (e.g. describe, discuss, identify) will help you to answer questions well in assessments.

Preparing for any assessment in VCE Chemistry requires organisational skills that can be practised throughout the year. These revision skills are very important in Units 3 and 4.

Organisational skills

The organisational skills that are useful for VCE studies include:

- creating a timetable for your studies, homework and other commitments
- setting SMART goals
- incorporating reflection and/or evaluation of strategies for continuous improvement.

Creating a timetable

It is important to have a balance when you are in VCE. Make sure you have time to study, continue your extracurricular activities (such as hobbies or a part-time job) and rest properly.

Follow the link in your obook pro to access a template you could use to create a study timetable, or you could use information technology resources to create your own version. Make sure you block out the time spent at school (including travel), playing sport, for family commitments, having meals/snacks and at part-time work. Be realistic – look at the amount of time you have left; you may have to make some sacrifices to succeed in your VCE studies without sacrificing your mental health.

1.14 Study plan template
Find me in your obook pro

For the remaining time, allocate homework and/or study time for each subject, making sure that you are actively applying the skills learnt in class rather than only trying to remember the content.

Setting SMART goals

Setting and achieving goals in VCE studies improves your motivation to succeed. Goals are most effective when they are SMART:

- Specific – clear and explicit
- Measurable – can be measured or is a measure of success
- Attainable – can be challenging (for learning and growth) but reachable
- Relevant – are meaningful and worthwhile (i.e. what is to be accomplished)
- Time-based – are committed to deadlines or set times.

When setting your SMART goals, consider your strengths and weaknesses, as well as your thoughts and feelings about subject(s). An example of a SMART goal would be ‘By the end of Term 1, I would like to achieve an average of B+ or higher for my internal assessment tasks in Chemistry’ rather than the less specific ‘improve in Chemistry’.

Reflection

Every few weeks, pause to reflect on your progress and the effectiveness of your organisational approach. Aim for continuous improvement – if something is not quite working, then change to a more reliable (and repeatable) strategy. Celebrate small successes and continue to challenge yourself by setting new goals.

The advantages of having good organisational skills include improving your marks in school-assessed coursework and reducing your stress, worry, **procrastination** and uncertainty levels. It will allow you to work smarter (not harder), avoid last-minute panics and gain a sense of accomplishment, in the knowledge that you did your best.

procrastination
the action of delaying or postponing something

Examination tips

When you sit your exam in Units 3 and 4, there are a few strategies that could help you prepare.

Before the exam:

- Start revising early. Cramming is stressful!
- Eat a healthy diet and cut back on processed or junk food and sugar-laden beverages.
- If you can, have a separate area where you revise or study from where you sleep because a good night's sleep is essential in the lead-up to the exam to help aid concentration.
- On the day of the examination, eat breakfast so that you are not distracted by hunger or the noise of your rumbling stomach.
- Take a bottle of water with the labels removed to maintain hydration.
- Arrive in plenty time before the examination; 15 minutes before the start of reading time is usually sufficient.
- Allow time for a bathroom visit before entering the examination venue.

During the exam:

- Use your reading time effectively – read the instructions and the questions carefully and plan out your responses.
- Once writing time has started, think about quickly jotting down some of the ideas that came to mind during reading time, so that you don't forget them later!
- Make use of your extra writing spaces or ask for a spare piece of paper.
- Don't dwell on questions you find difficult. Move on and return to them once you've finished the other questions well.
- On the multiple-choice questions, eliminate those answer choices that you think are incorrect.
- Beware of careless mistakes such as missing units, incorrect decimal places in conversions and using incorrect formulas and significant figures in answers.
- When drawing chemical compounds, chemical bond types or semi-structural formulas, use large, clear and accurate diagrams.

After you finish learning all of the content in Units 3 and 4, your teacher may assign you a practice exam that tests your knowledge and skills on everything you've learnt!

Study tip

Studying for your VCE exams should begin during the first week of the first topic, as smaller and consistent study sessions create better student outcomes (higher marks) than longer and irregular study sessions.

Understand command terms

Make sure you understand what the command terms mean, so that you can provide succinct and appropriate responses. Table 1 lists the VCAA VCE command terms that you may encounter in assessment tasks and exams.

TABLE 1 Command terms

Term	Explanation
account of	Describe a series of events or transactions.
account for	State reasons for; report on.
analyse	Identify components/elements and the significance of the relationship between them; draw out and relate implications; determine logic and reasonableness of information.
apply	Use; employ in a particular situation or context.
assess	Make a judgment about, or measure, determine or estimate, the value, quality, outcomes, results, size, significance, nature or extent of something.
calculate	Determine from given facts, figures or information; obtain a numerical answer showing the relevant stages in the working; determine or find (e.g. a number, answer) by using mathematical processes.
clarify	Make a statement or situation more comprehensible.
compare	Recognise similarities and differences and the significance of these similarities and differences.
construct	Make, build, create or put together by arranging ideas or items (e.g. an argument, artefact or solution); display information in a diagrammatic or logical form.
contrast	Show how things are different or opposite.
deduce	Draw a conclusion from given information, data, a narrative, an argument, an opinion, a design and/or a plan.
define	Give the precise meaning and identify essential qualities of a word, phrase, concept or physical quantity.
demonstrate	Show ideas, how something can be done or that something is true by using examples or practical applications, or by applying algorithms or formulas.
describe	Provide characteristics, features and qualities of a given concept, opinion, situation, event, process, effect, argument, narrative, text, experiment, artwork, performance piece or other artefact in an accurate way.
discuss	Present a clear, considered and balanced argument or prose that identifies issues and shows the strengths and weaknesses of, or points for and against, one or more arguments, concepts, factors, hypotheses, narratives and/or opinions.
distinguish	Make clear the differences between two or more arguments, concepts, opinions, narratives, artefacts, data points, trends and/or items.
evaluate	Ascertain the value or amount of; make a judgment using the information supplied, criteria and/or own knowledge and understanding to consider a logical argument and/or supporting evidence for and against different points, arguments, concepts, processes, opinions or other information.
examine	Consider an argument, concept, debate, data point, trend or artefact in a way that identifies assumptions, possibilities and interrelationships.
explain	Give a detailed account of why and/or how with reference to causes, effects, continuity, change, reasons or mechanisms; make the relationships between things evident.
extract	Select relevant and/or appropriate detail from an argument, issue or artefact.
extrapolate	Infer and/or extend information that may not be clearly stated from a narrative, opinion, graph or image by assuming existing trends will continue.
identify	Recognise and name and/or select an event, feature, ingredient, element, speaker and/or part from a list or extended narrative or argument, or within a diagram, structure, artwork or experiment.
infer	Derive conclusions from available information or evidence, or through reasoning, rather than through explicit statements.
interpret	Draw meaning from an argument, point of view, description or diagram, text, image or artwork and determine significance within context.
investigate	Observe, study or carry out an examination in order to establish facts and reach new conclusions.
justify	Show, prove or defend, with reasoning and evidence, an argument, decision and/or point of view using given data and/or other information.
list	Provide a series of related words, names, numbers or items that are arranged consecutively.

(*continued*)

TABLE 1 continued

Term	Explanation
name	Provide a word or term (something that is known and distinguished from other people or things) used to identify an object, person, thing, place etc.
outline	Provide an overview or the main features of an argument, point of view, text, narrative, diagram or image.
persuade	Induce (someone) to do something through reasoning or argument; convince.
predict	Give an expected result of an upcoming action or event; suggest what may happen based on available information.
propose	Suggest or put forward a point of view, idea, argument, diagram, plan and/or suggestion based on given data or stimulus material for consideration or action.
recall	Present remembered ideas, facts and/or experiences.
recommend	Put forward and/or approve (someone or something) as being suitable for a particular purpose or role.
recount	Retell a series of events or steps in a process, usually in order.
state	Give a specific name or value or other brief answer without explanation or calculation.
suggest	Put forward for consideration a solution, hypothesis, idea or other possible answer.
summarise	Retell concisely the relevant and major details of one or more arguments, text, narratives, methodologies, processes, outcomes and/or sequences of events.
synthesise	Combine various elements to make a whole or an overall point.

Source: *Glossary of command terms* reproduced by permission © VCAA

1.14 CHECK YOUR LEARNING

Describe and explain

1 Describe at least two strategies that can be useful for answering examination questions.

2 Explain why having good organisational skills is important for VCE studies.

Apply, analyse and compare

3 Analyse the revision strategies that you currently use and develop a SMART goal for improvement in this area.

4 Contrast the following command terms.

  a Deduce and calculate

  b Outline and identify

  c Evaluate and justify

Design and discuss

5 Use ICT resources to create and print your personal study timetable, using a template similar to that shown in Figure 1. Keep it visible and stick to it!

STUDY PLAN

Mon /	Tues /	Wed /	Thur /	Fri /	Sat /	Sun /
Day-	Day-	Day-	Day-	Day-	Day-	Day-
06:00						
07:00						
08:00						
09:00						
10:00						
11:00						
12:00						
13:00						
14:00						
15:00						
16:00						
17:00						
18:00						
19:00						
20:00						
21:00						
22:00						
23:00						
24:00						
To-do list						
•	•	•	•	•	•	•
•	•	•	•	•	•	•
•	•	•	•	•	•	•
•	•	•	•	•	•	•

FIGURE 1 An example of a study plan template

CHAPTER 1 Review

Chapter summary

1.1

- The VCE Chemistry course has a clear structure.
- There are many career pathways available from studying Chemistry.

1.2

- Aboriginal and Torres Strait Islander Peoples were Australia's first chemists and applied their unique scientific knowledge about the natural world to find (or refine) suitable food sources and develop traditional medicines and other useful materials in sustainable ways.

1.3

- The key science skills and their application are important for success in VCE Chemistry.
- A research question states the specific problem or issue on which your investigation will be based.
- An aim is a statement of what is to be investigated.
- A hypothesis is a testable statement that may include a prediction of an outcome of an investigation.

1.4

- A variety of methodologies can be used for a scientific investigation.
- A logbook is required for recording raw data and must be submitted intermittently for assessment purposes.
- The scientific poster has a specific format in VCE Science.

1.5

- Safety is your priority when performing scientific investigations.
- A risk assessment is an organised way of identifying hazards and/or risk factors and implementing controls for prevention when performing scientific investigations.

1.6

- Skills and knowledge of ethical understanding are applied to all types of scientific investigations.

1.7

- Data generated from scientific investigations is important evidence to support the trends, patterns and/or relationships between variables.
- Raw data can be qualitative or quantitative, discrete or continuous.
- Raw data from scientific investigations needs to be presented clearly so that it can be easily understood.

1.8

- Investigations that generate raw data must be valid, repeatable and reproducible.
- Errors and outliers must be included and accounted for in data evaluations.
- Significant figures must be considered in all calculations during VCE Chemistry.

1.9

- Analysis of raw primary data is important for constructing evidence-based arguments and conclusions.
- Sentence and paragraph structure is important for constructing evidence-based arguments and conclusions.

1.10

- Effective science communication depends on knowing your audience, avoiding jargon and prioritising information.

1.11

- The three perspectives of sustainability in VCE Chemistry are sustainable development, green chemistry principles, and linear and circular economies.

1.12

- The periodic table organises the elements in increasing order of atomic number. Each column of the periodic table is called a group and each row is called a period.
- Trends in the periodic table are explained by differences in number of electron shells (going down a group) and nuclear charge (going across a period).

1.13

- Balanced chemical equations obey the law of conservation of mass.
- Constructing and balancing chemical equations is a prerequisite for completing quantitative calculations in VCE Chemistry.

1.14

- Organisational skills are important for assessment preparation.
- Understanding command terms can help you to answer questions properly on assessment tasks.

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Mostly	Not really	Revision link
• understand the structure of the VCE Chemistry course, including the Areas of Study, their Outcomes, and Assessments	☐	☐	☐	Go back to Topic 1.1
• understand that Aboriginal and Torres Strait Islander Peoples' knowledges and perspectives have informed and enhanced scientific thinking	☐	☐	☐	Go back to Topic 1.2
• develop aims and questions, formulate hypotheses and make predictions	☐	☐	☐	Go back to Topic 1.3
• plan and conduct investigations, including using a logbook	☐	☐	☐	Go back to Topic 1.4
• describe the eight key scientific investigation methodologies relevant to VCE Chemistry	☐	☐	☐	Go back to Topic 1.4
• comply with safety and ethical guidelines, including understanding basic lab safety rules and the importance of risk assessments	☐	☐	☐	Go back to Topic 1.5
• describe the role of ethical understanding in VCE Chemistry	☐	☐	☐	Go back to Topic 1.6
• generate, collate and record data, including differentiating between different types of data (qualitative, quantitative, discrete and continuous data) and using graphs to represent data	☐	☐	☐	Go back to Topic 1.7
• analyse and evaluate data and investigation methods, including considering the validity, repeatability and reproducibility of data, identifying errors and outliers, and using significant figures	☐	☐	☐	Go back to Topic 1.8
• construct evidence-based arguments and draw conclusions	☐	☐	☐	Go back to Topic 1.9
• analyse, evaluate and communicate scientific ideas	☐	☐	☐	Go back to Topic 1.10
• describe the three different perspectives on sustainability, including sustainable development, green chemistry principles, and linear and circular economies	☐	☐	☐	Go back to Topic 1.11
• use the periodic table as a tool to identify patterns and properties of elements	☐	☐	☐	Go back to Topic 1.12
• balance chemical equations	☐	☐	☐	Go back to Topic 1.13
• understand the importance of organisational skills in preparing for assessments, including creating and using a timetable, setting SMART goals and reflecting on progress	☐	☐	☐	Go back to Topic 1.14

Revision questions

Multiple choice

1 The law of conservation of mass means that:

A chemical formulas are conserved during a reaction.

B there are different types of atoms in the reactants from those of the products.

C the coefficients used to balance chemical equations.

D the mass of the reactants must equal the mass of the products.

2 The balanced chemical equation for the burning of propane in oxygen is:

A $C_3H_8 + 10O_2 \rightarrow 3CO_2 + 4H_2O$

B $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

C $C_3H_8 + 13O_2 \rightarrow 3CO_2 + 5H_2O$

D $2C_3H_8 + 13O_2 \rightarrow 6CO_2 + 10H_2O$

3 Which of the following best describes the required VCAA format for the scientific poster?

A Aim, hypothesis, methodology, results, discussion, conclusion, references

B Introduction, methodology, results, communication statement, discussion, conclusion, references

C Title, student name, introduction, methodology, results, communication statement, discussion, conclusion, references and acknowledgements

D Abstract, introduction, methodology, results, discussion, conclusion, acknowledgements

4 What is the variable that is purposely changed in an experiment or practical activity?

A Controlled variable

B Dependent variable

C Independent variable

D Random variable

5 How many significant figures are in the measurement: Temperature = 0.0230°C?

A 2

B 3

C 4

D 5

6 Poor precision in scientific measurements may arise from:

A theoretical values being too inflexible.

B human error.

C limitations of the measuring instrument.

D both human error and the limitations of the measuring instrument.

7 If an accepted value is 15.63 and a student measured 12.82, 12.96, 13.02 and 12.99 in subsequent trials of an experiment, their data would be described as:

A accurate but not precise.

B precise but not accurate.

C both accurate and precise.

D neither accurate nor precise.

8 Deduce which of the following is not compatible with the principles of green chemistry.

A Maximising energy use in each process

B Minimising the use of toxic chemicals by replacing them with safer alternatives

C Maximising the atom efficiency of each reaction pathway

D Minimising the formation of wastes and by-products

Short answer

Describe and explain

9 Describe at least two examples of chemical processes used by Aboriginal or Torres Strait Islander Peoples for food sources, medicinal or construction purposes that are not mentioned in Topic 1.2.

10 Explain some of the language features used for constructing evidence-based arguments for scientific investigations.

Apply, analyse and compare

11 A student set up an electrochemical cell (Figure 1) to investigate the effect of temperature on measured cell voltage. Solutions of the electrolytes used were 1.0 M. The results of the practical investigation are given in Table 1.

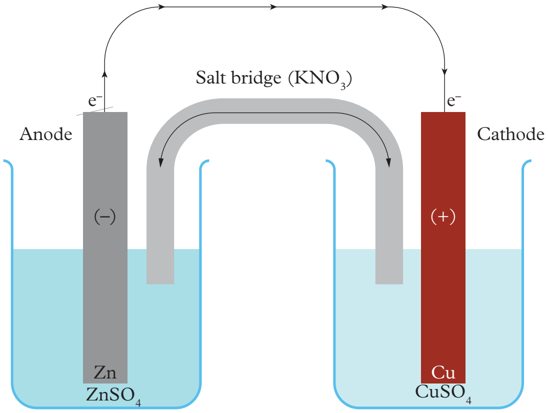


FIGURE 1 Electrochemical cell set-up for the temperature–voltage practical investigation

a Identify the dependent and the independent variables in this practical investigation.

b Write an aim for this practical investigation.

c Formulate a hypothesis for this practical investigation.

d Calculate the average voltage reading for each temperature recorded and complete Table 1.

e Create a visual representation of the data in Table 1 (with appropriate labels) to identify the relationship between the dependent and independent variables in this experiment.

f Use the electrochemical series (page 135) to calculate the percentage error of the cell voltage predicted at 25°C. Express your answer to 2 significant figures.

Design and discuss

12 Design an evidence-based analysis and conclusion for the data presented in Table 1 and the visual representation from Question 11e.

13 Design a risk assessment for the practical investigation in Question 11.

TABLE 1 Results of the practical investigation

Temperature (°C)	Voltage					
	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5	Average
0	0.91	0.90	0.91	0.87	0.89	
3	0.91	0.94	0.90	0.95	0.93	
15	0.83	0.85	0.86	0.83	0.86	
22	0.84	0.85	0.82	0.86	0.85	
54	0.85	0.82	0.78	0.85	0.84	

You can find the following resources for this section in your obook pro:

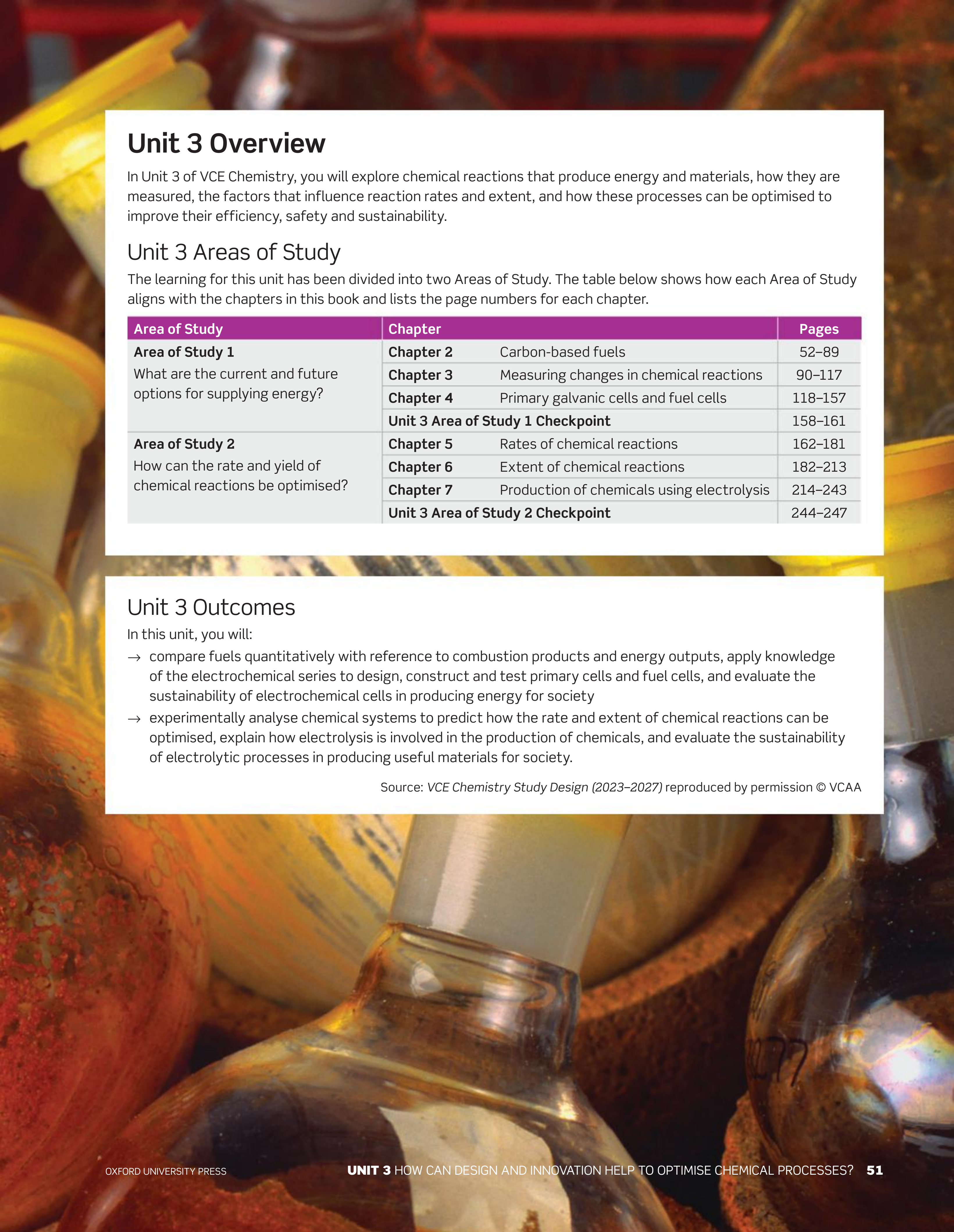
Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

UNIT 3

HOW CAN DESIGN AND INNOVATION HELP TO OPTIMISE CHEMICAL PROCESSES?

FIGURE 1 Chemical processes could be improved to make use of residues from reactions.



Unit 3 Overview

In Unit 3 of VCE Chemistry, you will explore chemical reactions that produce energy and materials, how they are measured, the factors that influence reaction rates and extent, and how these processes can be optimised to improve their efficiency, safety and sustainability.

Unit 3 Areas of Study

The learning for this unit has been divided into two Areas of Study. The table below shows how each Area of Study aligns with the chapters in this book and lists the page numbers for each chapter.

Area of Study	Chapter		Pages
Area of Study 1 What are the current and future options for supplying energy?	**Chapter 2**	Carbon-based fuels	52–89
	Chapter 3	Measuring changes in chemical reactions	90–117
	Chapter 4	Primary galvanic cells and fuel cells	118–157
	Unit 3 Area of Study 1 Checkpoint		158–161
Area of Study 2 How can the rate and yield of chemical reactions be optimised?	**Chapter 5**	Rates of chemical reactions	162–181
	Chapter 6	Extent of chemical reactions	182–213
	Chapter 7	Production of chemicals using electrolysis	214–243
	Unit 3 Area of Study 2 Checkpoint		244–247

Unit 3 Outcomes

In this unit, you will:

- → compare fuels quantitatively with reference to combustion products and energy outputs, apply knowledge of the electrochemical series to design, construct and test primary cells and fuel cells, and evaluate the sustainability of electrochemical cells in producing energy for society
- → experimentally analyse chemical systems to predict how the rate and extent of chemical reactions can be optimised, explain how electrolysis is involved in the production of chemicals, and evaluate the sustainability of electrolytic processes in producing useful materials for society.

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CHAPTER 2

Carbon-based fuels

KEY KNOWLEDGE

- → the definition of a fuel, including the distinction between fossil fuels (coal, natural gas, petrol) and biofuels (biogas, bioethanol, biodiesel) with reference to their renewability (ability of a resource to be replaced by natural processes within a relatively short period of time)
- → fuel sources for the body measured in kJ g^{-1}: carbohydrates, proteins and lipids (fats and oils)
- → photosynthesis as the process that converts light energy into chemical energy and as a source of glucose and oxygen for respiration in living things:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

- → oxidation of glucose as the primary carbohydrate energy source including the balanced equation for cellular respiration:

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

- → production of bioethanol by the fermentation of glucose:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$$

  and subsequent distillation to produce a more sustainable transport fuel
- → comparison of exothermic and endothermic reactions, with reference to bond making and bond breaking, including enthalpy changes (ΔH) measured in kJ, molar enthalpy changes measured in kJ mol^{-1} and enthalpy changes for mixtures measured in kJ g^{-1} and their representations in energy profile diagrams
- → determination of limiting reactants or reagents in chemical reactions
- → combustion (complete and incomplete) reactions of fuels as exothermic reactions: the writing of balanced thermochemical equations, including states, for the complete and incomplete combustion of organic molecules using experimental data and data tables.

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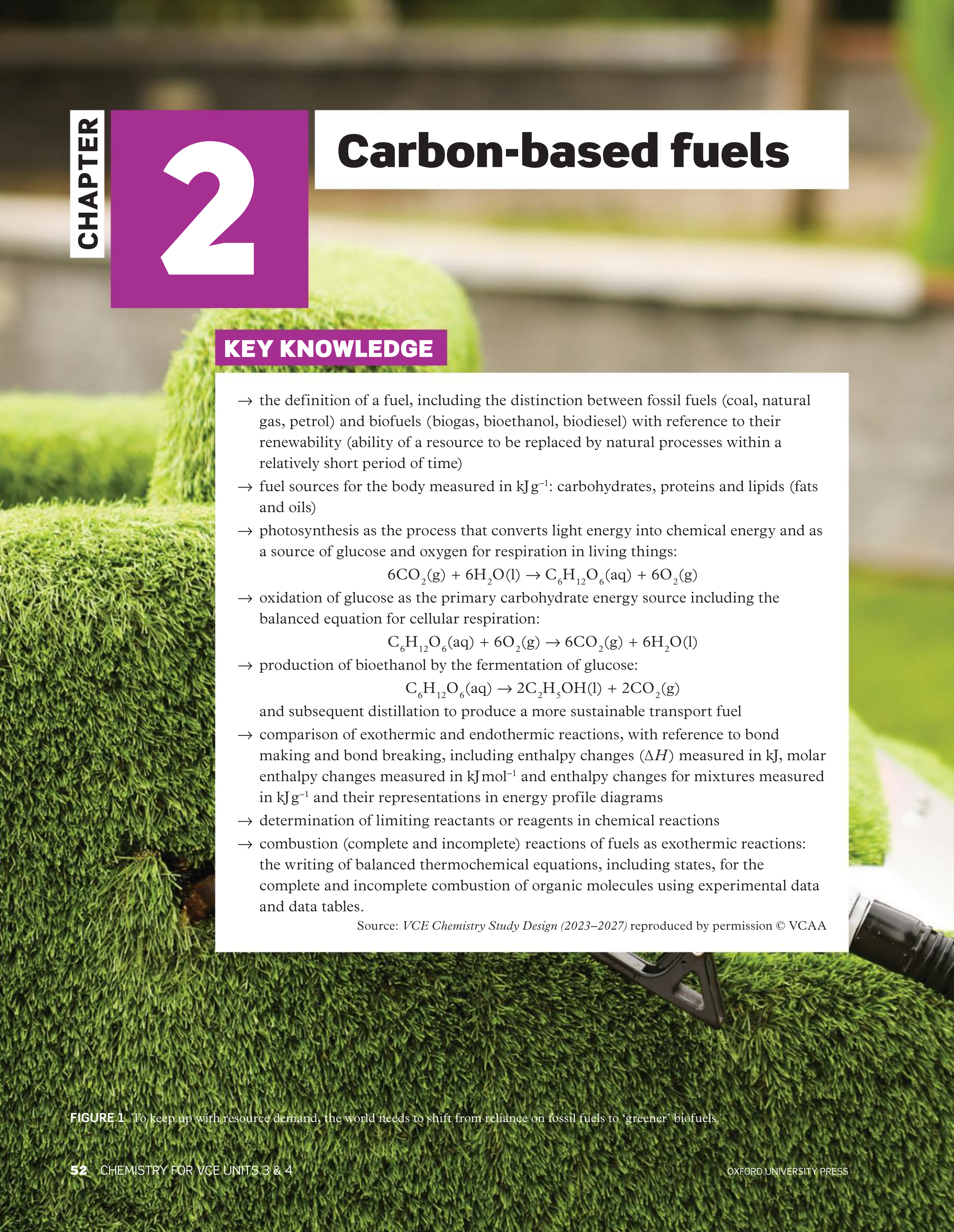
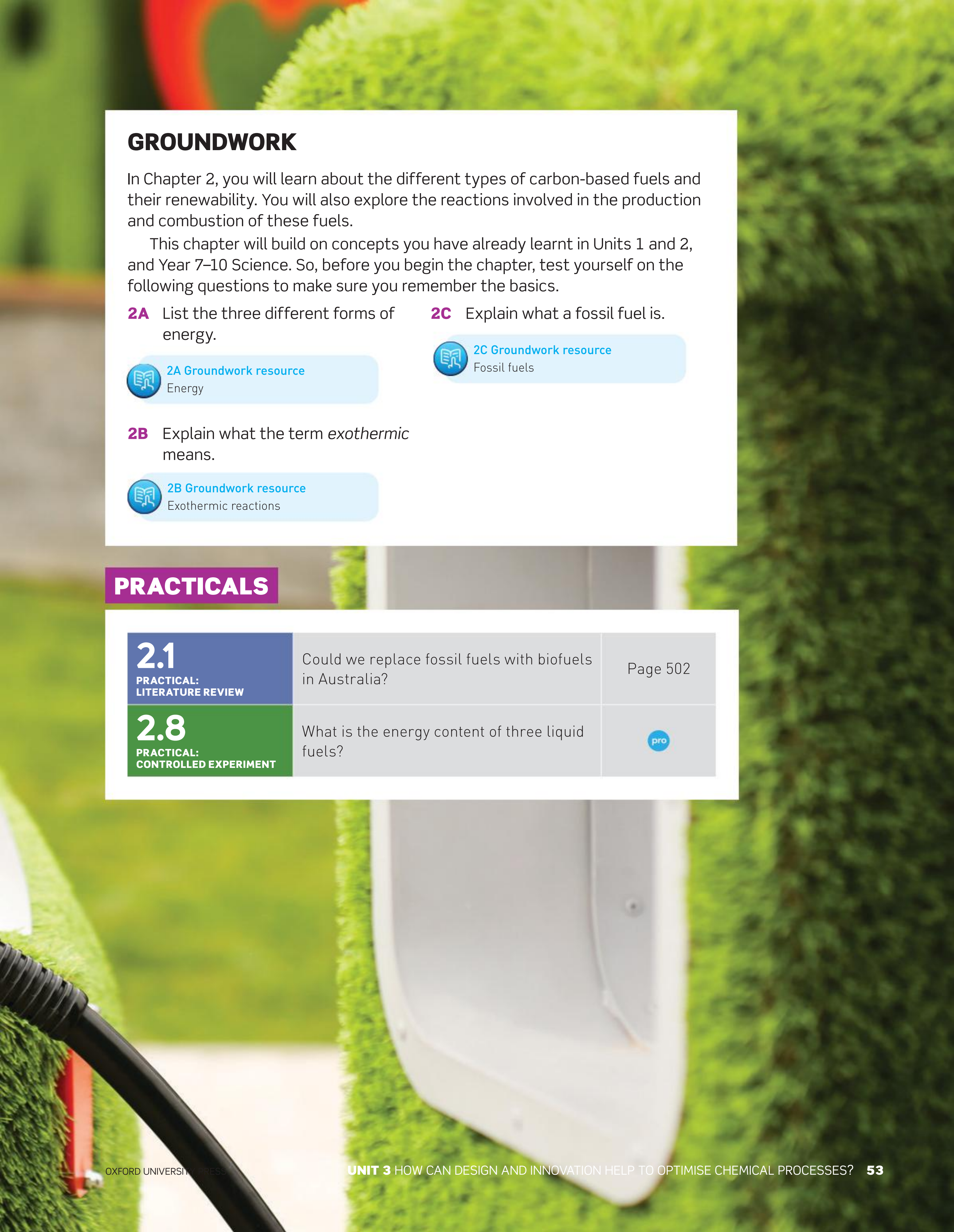


FIGURE 1 To keep up with resource demand, the world needs to shift from reliance on fossil fuels to 'greener' biofuels.



GROUNDWORK

In Chapter 2, you will learn about the different types of carbon-based fuels and their renewability. You will also explore the reactions involved in the production and combustion of these fuels.

This chapter will build on concepts you have already learnt in Units 1 and 2, and Year 7–10 Science. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

2A List the three different forms of energy.

2A Groundwork resource
Energy

2B Explain what the term *exothermic* means.

2B Groundwork resource
Exothermic reactions

2C Explain what a fossil fuel is.

2C Groundwork resource
Fossil fuels

PRACTICALS

2.1 **PRACTICAL: LITERATURE REVIEW**	Could we replace fossil fuels with biofuels in Australia?	Page 502
2.8 **PRACTICAL: CONTROLLED EXPERIMENT**	What is the energy content of three liquid fuels?	pro

2.1

The renewability of fuels

KEY IDEAS

In this topic, you will learn that:

- a fuel is a substance that can be reacted with another substance (e.g. combusted with oxygen in the air) to produce useful energy
- fossil fuels are non-renewable mixtures of hydrocarbons extracted from the ground, such as coal, oil and natural gas
- biofuels are renewable fuels made from plants that can be replaced by natural processes within a relatively short period of time.

fuel	a substance that reacts to produce heat energy
combustion	a reaction with oxygen that produces carbon dioxide, water and heat energy
renewable	replenishable by natural processes in a relatively short amount of time
fossil fuel	a non-renewable fuel, such as coal, crude oil and natural gas, obtained from decomposed organic matter that has been compressed under heat and pressure for millions of years underground
non-renewable	takes a very long time to form; is used more quickly than it can be produced
biofuel	fuels produced from plants
petrodiesel	diesel fuel obtained from crude oil; a fossil fuel

A **fuel** is a substance that provides heat energy when it is reacted with another substance. Fuels include petrol, diesel, propane, methanol, coal, wood and biodiesel. They are usually compounds of carbon, hydrogen and sometimes oxygen.

Combustion reactions between these compounds and oxygen in the air release heat (or thermal) energy, which can be harnessed in an engine for useful work, such as to turn the wheels of a road vehicle or the turbine of an aeroplane. All combustion reactions of fuels are considered exothermic, because they release heat energy.

Renewability of fuels

Fuels are defined as **renewable** when they can be replaced by natural processes within a relatively short period of time. **Fossil fuels** such as coal, oil and natural gas are therefore **non-renewable** because they take millions of years to form. **Biofuels** are renewable because they are made from plant material, which can be regrown in a matter of years or less.

Fossil fuels

As phytoplankton and tiny marine animals die, they sink to the ocean floor and become compressed over millions of years by the pressure of the ocean above (Figure 1), forming oil and gas. Many large fossil fuel deposits we know of today formed underwater many hundreds of millions of years ago. Central Australia's large oil deposits formed when the area was submerged under ocean during the Cretaceous period, around 100 million years ago, and they are an important source of **petrodiesel** today.

Coal is formed by plants that grow on land, die and decay, and then turn into peat; the peat turns into coal over millions of years. Fossil-based hydrocarbon fuels (made up of hydrogen and carbon molecules) are therefore not considered to be renewable. We use them at a much faster rate than we can replace them.

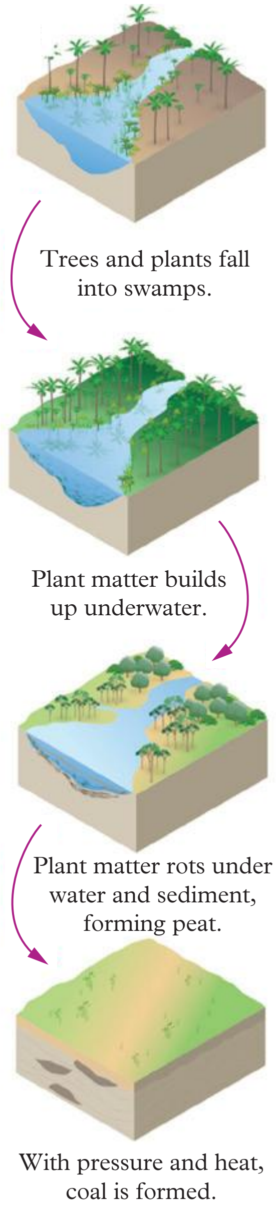


FIGURE 1 The formation of fossil fuels, such as coal, takes millions of years.

2.1
PRACTICAL: LITERATURE REVIEW
Could we replace fossil fuels with biofuels in Australia? Go to page **502**.

Biofuels

Biofuels are renewable because they are made from easily regenerable organic substances like plants. Fast-growing plant species which produce high **yields** are specifically grown for the purpose of biofuel production. Selected plants also efficiently convert energy from the Sun's light into chemical energy. The plants' chemical energy can then be converted to chemical energy in the fuel.

yield
the total amount of product formed from a chemical reaction

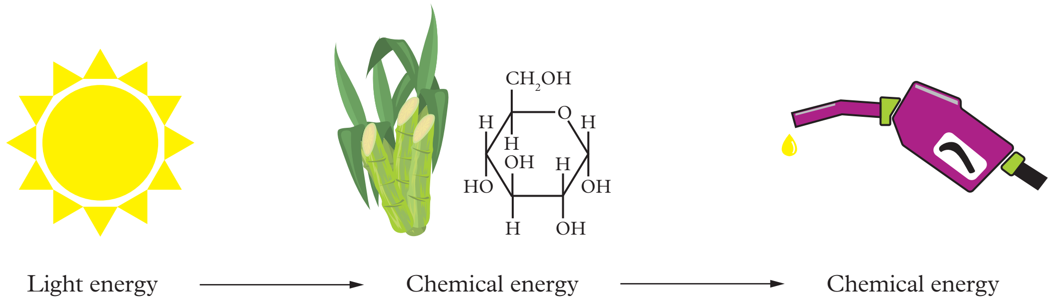


FIGURE 2 The energy transfers that occur in biofuel production

Let's explore three types of biofuels: biogas, bioethanol and biodiesel.

Biogas

Biogas is a mixture of gases, including methane, carbon dioxide and other gases. It is produced by the anaerobic fermentation of vegetable waste by bacteria. Biogas reactors work in a very similar way to vegetable composters, except that the gas is collected in a storage tank.

As biogas contains significant amounts of carbon dioxide and other impurities that could damage engines, it is commonly combusted for heating and tends to be used on-site (e.g. at a farm) instead of being sold on the open market. The advantages of using biogas are:

- being made from plants, biogas is renewable and reduces our reliance on fossil fuels
- collecting and combusting biogas results in lower greenhouse gas emissions than allowing the methane, which is a powerful greenhouse gas, to escape to the atmosphere.

biogas
a fuel consisting of a blend of gases including methane and carbon dioxide, produced by the anaerobic fermentation of plant waste in a biogas reactor vessel

Bioethanol

Bioethanol is ethanol produced from the anaerobic fermentation of glucose by **enzymes** according to the following equation:

$$C_6H_{12}O_6(aq) \rightarrow 2CH_3CH_2OH(l) + 2CO_2(g)$$

The glucose used in this reaction is sourced from plants. It can therefore be replaced by natural processes within a relatively short period of time and is considered renewable.

When ethanol is mixed with petrol to give a final concentration of no more than 10%, it can be used in almost all petrol engines. This fuel is called **E10** and can help slightly decrease our reliance on fossil fuels. The production, advantages and disadvantages of bioethanol are explored in more detail in Topic 2.5.

FIGURE 3 The E10 petrol offered at petrol stations contains up to 10% ethanol.

bioethanol
ethanol produced by the fermentation of sugars (e.g. glucose) by yeast enzymes, then distilled to a high enough purity (>95%) to be used in an internal combustion engine

enzyme
a protein-based substance produced by living organisms; speeds up chemical reactions

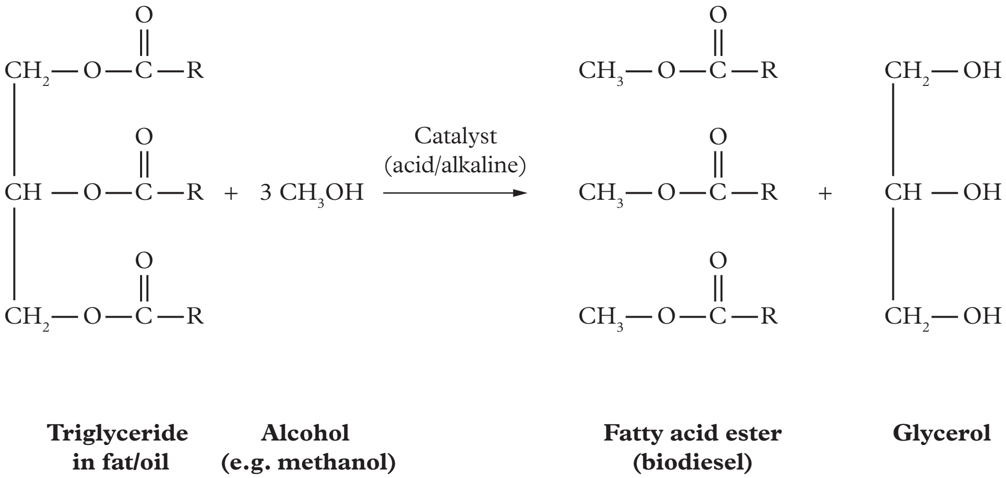
E10
a fuel blend of up to 10% bioethanol and 90% or more of petrol

Biodiesel

biodiesel
a renewable fuel consisting of a mixture of fatty acid methyl esters produced by the transesterification reaction between plant oils and methanol or ethanol

Biodiesels are a blend of fatty acids that have been synthesised via a transesterification reaction from compounds found in plant oils called triglycerides. The equation for the formation of biodiesel is shown below. You will learn about this reaction in greater detail in Chapter 9.

Oils from canola, corn, soybean, grapeseed, and sunflower are just some of the common sources of biodiesels.



$$\text{triglyceride} + 3\,CH_3OH \xrightarrow[\text{(acid/alkaline)}]{\text{Catalyst}} 3\,CH_3-O-\overset{\overset{\displaystyle O}{\|}}{C}-R + CH_2(OH)-CH(OH)-CH_2(OH)$$

FIGURE 4 Biodiesel is produced by a transesterification reaction.

Biodiesel offers two main advantages over petrodiesel (from fossil fuels). Biodiesel is made from plant material and is therefore considered renewable. The plant material can be replenished in a relatively short period of time. As biodiesel is made from plant oils, it also helps reduce our reliance on fossil fuels.

The properties of biodiesels are slightly different to those of petrodiesel. Biodiesel is more viscous, has a slightly lower energy content and has a much higher melting point. This can make it unsuitable for use in cold climates because it can 'gel' or even freeze during cold weather, and damage engines. For these reasons, biodiesel is usually blended into existing petrodiesel fuel at concentrations of no more than 20% to avoid engine modifications. This results in a fuel called **B20**.

B20
a fuel blend of less than 20% biodiesel and more than 80% petrodiesel

FIGURE 5 Sunflower oil is a common source of biodiesel.

2.1 REAL-WORLD CHEMISTRY

Biofuels in Brazil

Climate, fertile land and government policies favouring biofuel usage have made Brazil the world's leading exporter of bioethanol. Domestic consumption of bioethanol in Brazil is also very high due to government policies implemented since the 1970s that encouraged the production and usage of bioethanol in cars. Here, sugar cane is used as the primary raw material for the synthesis of bioethanol.

FIGURE 6 Sugar cane and ethanol mill in Brazil

Apply your understanding

1 List two advantages of adding bioethanol to petrol.
2 Suggest two reasons why other countries may be unable to compete with Brazil at producing bioethanol.
3 Suggest two disadvantages of bioethanol production using sugar cane in Brazil.

2.1 CHECK YOUR LEARNING

Describe and explain

1 Define the term *fuel*.
2 Describe the process of producing biogas.
3 Outline one advantage of using biodiesel over petrodiesel.

Apply, analyse and compare

4 Distinguish between fossil fuels and biofuels.
5 Summarise the differences between biodiesel and petrodiesel.

Design and discuss

6 Discuss any advantages that are common to all three biofuels described in this section.
7 Discuss any disadvantages that are common to all three biofuels described in this section.

FIGURE 7 Biofuel

2.2 Fuel sources for the body

KEY IDEAS

In this topic, you will learn that:

- carbohydrates, proteins and lipids (fats and oils) are fuel sources for the human body
- fuel sources for the body are measured in $kJ\,g^{-1}$.

Carbohydrates, proteins and lipids

Our bodies require nutrition in the form of **macronutrients** such as **carbohydrates**, **proteins** and **lipids**, as well as other substances. These large molecules (macromolecules) must first be broken down by the digestive system before they can be used by the body's cells. Macronutrients mainly consist of carbon, hydrogen, oxygen and nitrogen.

macronutrient
type of nutrient needed by the body in large amounts; macromolecules that provide energy to the body

carbohydrate
a macronutrient consisting of one or more sugar units

protein
a macronutrient consisting of a chain of amino acids

lipid
a macronutrient that has a very low solubility in water

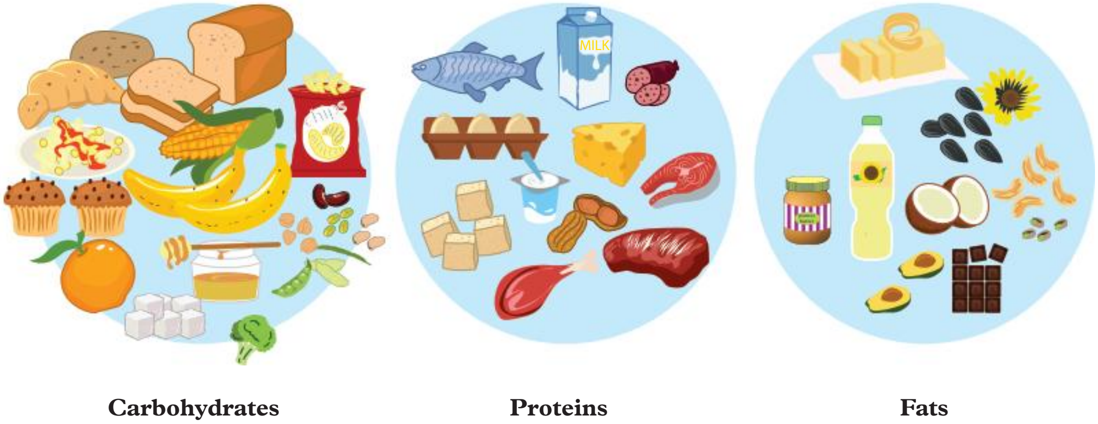


FIGURE 1 Common dietary sources of the three macromolecules: carbohydrates, proteins and fats

In our bodies, cellular respiration combines these nutrients with oxygen to release energy in an exothermic process. You will learn more about this process in Topic 2.4. This energy is essential for metabolic functions in the body. Each of the macronutrients provides different amounts of energy per gram ingested.

When a macronutrient is combusted, a certain amount of energy is released. The total energy available to be used by the body is always slightly lower than what is released by combustion. This is because the body's enzymatic processes are not able to completely oxidise and absorb all nutrients. **Energy content** is the amount of energy a macronutrient provides that can be used by the body.

energy content
the total energy released when nutrient molecules are combusted, which can be used by the body

Carbohydrates

Carbohydrates are found in a range of different foods. They have a variety of roles in living organisms, including energy storage and as building blocks for cells. Carbohydrates are made from sugar units or monomers, called **saccharides**. Large carbohydrates that consist of more than three saccharides are called **polysaccharides**. Carbohydrates such as glucose, sucrose and starch have varying energy contents per gram. Of the approximately $17\,kJ\,g^{-1}$ that could be obtained from combusting carbohydrates, around $16\,kJ\,g^{-1}$ can be recovered by the body's metabolic processes. The energy content of carbohydrates is therefore $16\,kJ\,g^{-1}$.

saccharides
sugars

polysaccharides
long-chain molecules made up of multiple saccharides

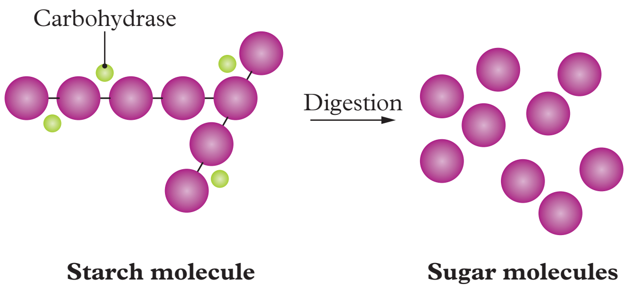


FIGURE 2 During digestion, the polysaccharide starch is broken down into smaller sugars by carbohydrases (enzymes that break down carbohydrates).

Explore the properties of a class of carbohydrates called steviol glycosides in Real-world chemistry 2.2.

2.2 REAL-WORLD CHEMISTRY

Steviol glycosides in fizzy drinks

Over the last decade, consumer demand for low-calorie sweeteners has increased. These sweeteners have primarily been derived from plant sources. Steviol glycosides, one of the most popular sweeteners, are extracted from the *Stevia rebaudiana* plant, commonly known as sweetleaf or candyleaf.

Steviol glycosides (also listed as additive 960) provide a similar energy content to digestible sugars such as glucose and sucrose, but taste up to 300 times sweeter. This allows manufacturers to achieve the desired level of sweetness in food using very little sweetener. Steviol glycosides are an approved food additive in almost all countries and are generally recognised as safe for consumption by children and adults.

Apply your understanding

1. Research the ingredients of two popular sweetened beverages such as cola and lemonade. List the sweeteners in the drink.
2. Research the ingredients of one popular low-calorie beverage such as cola or lemonade. Compare the sweeteners in this drink with the sweeteners in the regular drink from Question 1.

Proteins

Proteins are long chains of **amino acids**. They are required for growth and repair of body tissues. Proteins can also be metabolised to produce useful energy if excessive amounts are ingested or, more commonly, when lipid or carbohydrate reserves are severely low. The process of extracting energy from proteins is less efficient than that for carbohydrates. Approximately $24\,\mathrm{kJ\,g^{-1}}$ is produced during combustion of a protein, but only around $17\,\mathrm{kJ\,g^{-1}}$ can be used by the body.

amino acid
a small organic compound which contains a $-NH_2$ and $-COOH$ functional group

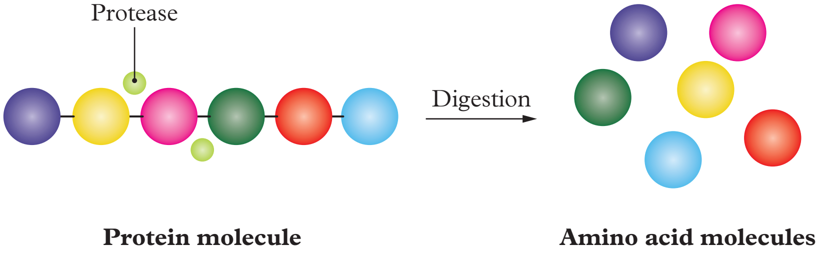


FIGURE 3 During digestion, protein molecules are broken down into amino acids by proteases (enzymes that break down proteins).

Lipids

Lipids are important building blocks of all living cells. They are a primary component of the cell membrane and are involved in a range of functions such as hormone regulation, nervous system function and energy storage. Examples include cholesterol and triglycerides. Lipids are characterised by their poor water solubility, but high solubility in organic solvents. They are categorised broadly as either fats or oils, depending on their state at 25°C. Fats are usually made by animals and are commonly solid at 25°C. Oils are usually made by plants and are liquid at 25°C. Despite their differing sources, fats and oils are chemically very similar to each other. When combusted, the energy released from lipids is 39 kJ g^{-1}, while the energy content is approximately 37 kJ g^{-1}.

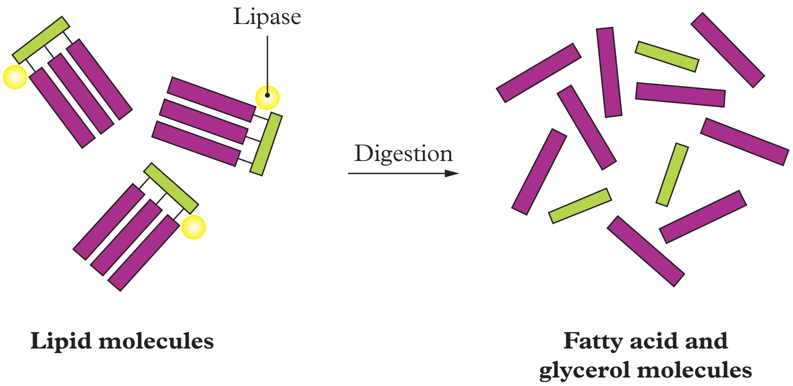


FIGURE 4 During digestion, lipid molecules are broken down into fatty acids and glycerol by lipases (enzymes that break down lipids).

Oxidation of nutrients to produce energy

The products of digestion can be oxidised by the body's cells to produce energy or to be stored for later energy use. The equation for the complete oxidation of a nutrient such as glucose or a fatty acid is the same as that for the combustion of that substance. When writing the equation, the carbon atoms should be balanced first, followed by hydrogen and, finally, oxygen.

For example:

$$CH_3(CH_2)_{14}COOH + 23O_2 \rightarrow 16CO_2 + 16H_2O$$

$$\text{palmitic acid} + \text{oxygen} \rightarrow \text{carbon dioxide gas} + \text{water}$$

2.2 CHALLENGE

Energy content from bagels

Calculate the energy content per 100 g of sesame bagels, in kJ, given that each 105 g sesame bagel contains 11.0 g of protein, 7.5 g of fat and 22.5 g of carbohydrate.

2.2 WORKED EXAMPLE

CHECKING THE ENERGY CONTENT ON NUTRITION LABELS

Nutrition labels should provide accurate information about the energy content in food provided by carbohydrates, fats and proteins. Here is a simplified nutrition label for a 100 g serving of muesli:

Energy content: 1728 kJ	
Carbohydrates: 57.8 g	Fat: 13.1 g
Sodium: 124.4 mg	Protein: 12.0 g

Calculate the amount of energy, in kJ, provided by:

a the fat content in 100 g muesli

b the carbohydrate content in 100 g muesli

c the protein content in 100 g muesli.

Think	Do
Step 1: Write down the available energy for each macronutrient type.	Lipids: 37 kJ g^{-1} Carbohydrates: 16 kJ g^{-1} Proteins: 17 kJ g^{-1}
Step 2: Calculate the amount of energy provided by the fat content by multiplying the grams of fat in the sample by the energy content of fat.	**a** 13.1 g × 37 kJ g^{-1} = 485 kJ (3 sig fig)
Step 3: Calculate the amount of energy provided by the carbohydrate content by multiplying the grams of carbohydrate in the sample by the energy content of carbohydrates.	**b** 57.8 g × 16 kJ g^{-1} = 925 kJ (3 sig fig)
Step 4: Calculate the amount of energy provided by the protein content by multiplying the grams of protein in the sample by the energy content of proteins.	**c** 12.0 g × 17 kJ g^{-1} = 204 kJ ≈ 200 kJ or 2.0 × 10^2 (2 sig fig)

2.2 CHECK YOUR LEARNING

Describe and explain

1 State the energy content, in kJ g^{-1}, for the following macronutrients.

- **a** Carbohydrates
- **b** Proteins
- **c** Fats

Apply, analyse and compare

2 Compare the difference between energy content and energy released on combustion.

3 Write a chemical reaction equation for the complete oxidation of linoleic acid.

4 Calculate the mass of fat contained within one 170 g avocado using the following nutritional information.

Nutrient	Value per 100 g of avocado (g)
Fat	15.0
Carbohydrate	9.0
Protein	2.0

5 Using the nutritional information from Question 4, calculate the total amount of energy provided to the human body by eating one 170 g avocado.

Design and discuss

6 The energy content values (in kJ g^{-1}) of carbohydrates, proteins and lipids are approximate because different types of carbohydrates, proteins and lipids have slightly different energy content values. Suggest one reason why different types of lipids may have different energy content values.

2.3 Photosynthesis

KEY IDEAS

In this topic, you will learn that:

- photosynthesis is the process that converts light energy into chemical energy
- the chemical equation for photosynthesis is:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

Photosynthesis

chloroplast the structure in living cells in which photosynthesis occurs

photosynthesis the process in which light energy is converted to chemical energy

Plants and other organisms that contain **chloroplasts** can obtain energy by capturing carbon dioxide from the atmosphere and converting it into the carbohydrate glucose. These organisms can then use chemical energy from glucose for their metabolic processes. Glucose can be further modified to make a wide variety of other molecules required by the plant or organism for its survival. The reaction that facilitates the conversion of light energy into chemical energy is called **photosynthesis**.

Energy conversions in photosynthesis

Photosynthesis is an endothermic reaction because it requires the absorption of light energy from the Sun. This is then converted into chemical energy by the plant's highly specialised enzymes (specific proteins that facilitate and speed up reactions). The enzymes capture carbon dioxide and convert it into glucose. Since photosynthesis relies on light energy, plants cannot capture carbon dioxide at night, or in the dark. Plants obtain energy at night via cellular respiration.

Chemical equation for photosynthesis

The reactants of photosynthesis are carbon dioxide, water and sunlight and the products are glucose and oxygen. The reaction is described using the following chemical equation:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

$$\text{carbon dioxide} + \text{water} \rightarrow \text{glucose} + \text{oxygen}$$

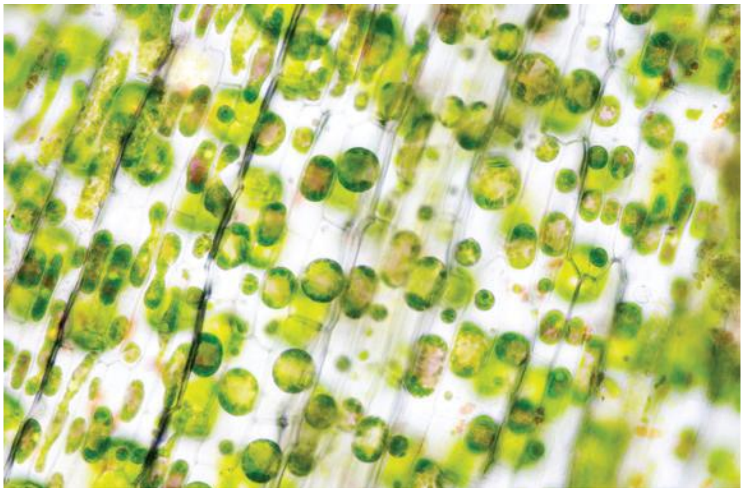


FIGURE 1 Photosynthesis occurs in the chloroplasts (green).

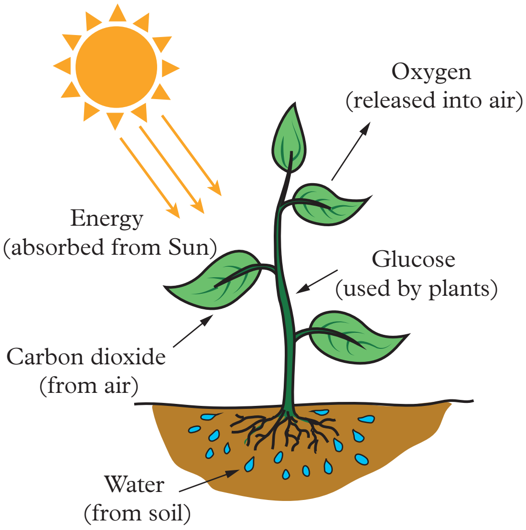


FIGURE 2 The reactants and products of photosynthesis

FIGURE 3
Photosynthesis converts light energy from the Sun into chemical energy that can be used by plants.

Efficiency of photosynthesis

Photosynthesis is not a very efficient process. In a typical crop plant (e.g. wheat), less than 1% of the energy from incoming sunlight is captured and used successfully by the plant's enzymes, and converted into chemical energy as glucose. For a non-crop plant, the efficiency can be around 0.2%. The highest energy conversion efficiencies have been observed in sugar cane and pineapples at around 3–4%. However, this is still much lower than the energy conversion efficiencies in non-living systems such as internal combustion engines (~40%), LED lights (~85%), and even photovoltaic cells (solar panels, ~15%).

Factors affecting the rate of photosynthesis

The rate of photosynthesis is largely determined by three factors. The first is light intensity. More incoming light increases the rate of photosynthesis. However, the rate plateaus in abundant sunlight. The second factor is carbon dioxide concentration. Like light intensity, a higher concentration of atmospheric carbon dioxide increases the rate at which plants can photosynthesise, but the rate plateaus once carbon dioxide is no longer a limiting factor. Finally, the rate of photosynthesis is at a maximum when the temperature is optimal (~20–25°C).

2.3 SKILL DRILL

Evaluating the ethics of genetically modifying crops

Key science skill: Analyse, evaluate and communicate scientific ideas

Realizing Increased Photosynthetic Efficiency (RIPE) is an international research project that aims to increase the efficiency of photosynthesis using genetic engineering. RIPE's research teams look for genetic modifications that can improve the efficiency of natural photosynthesis, then transfer these genes into cowpea, cassava, rice, and soybean crops. If successful, RIPE could make genetic modifications to existing crops to make them grow faster and produce more food.

Practise your skills

1. Suggest one advantage of improving the photosynthetic efficiency of rice crops.
2. In addition to plant genetics, list two other factors that can affect the rate of efficiency of photosynthesis.
3. Genetically modified organisms are often discussed in ethical debates. Research the advantages and disadvantages of genetically modifying plants and discuss whether you believe improving the photosynthetic efficiency of rice crops is justified.

Need help analysing, evaluating or communicating scientific ideas? See Topics 1.8–1.10 (pages 24–31).

2.3 CHECK YOUR LEARNING

Describe and explain

1. Write a balanced chemical equation for photosynthesis in plants.
2. List three factors that affect the rate of photosynthesis in plants.
3. Explain whether photosynthesis is an endothermic or exothermic reaction.

Apply, analyse and compare

4. Compare the photosynthetic efficiency of rice, grass and pineapples.

Design and discuss

5. Suggest why the photosynthetic efficiencies of food crops such as wheat are generally higher than of non-food plants such as pine trees.

2.4

Oxidation of glucose

KEY IDEAS

In this topic, you will learn that:

- glucose is the main carbohydrate used as an energy source by plants and animals
- the balanced equation for cellular respiration is:

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

adenosine triphosphate (ATP)
the organic molecule that provides energy for metabolic processes in an organism

cellular respiration
a combustion reaction in which nutrient molecules react with oxygen to produce energy

aerobic respiration
cellular respiration in the presence of oxygen

anaerobic fermentation
cellular respiration in the absence of oxygen

monosaccharides
molecules made up of single saccharide units

Cellular respiration in plants and animals

Glucose is used by plants and animals as a source of chemical energy, or fuel. Once glucose has been obtained or synthesised, a series of reactions occurs to produce chemical energy in the form of **adenosine triphosphate** (ATP). This process is called **cellular respiration**.

There are two types of cellular respiration:

- **Aerobic respiration** – requires oxygen and yields a high amount of energy in the form of ATP.
- **Anaerobic fermentation** – does not require oxygen and yields less ATP than aerobic respiration.

You will learn about anaerobic fermentation in the next topic.

Aerobic cellular respiration

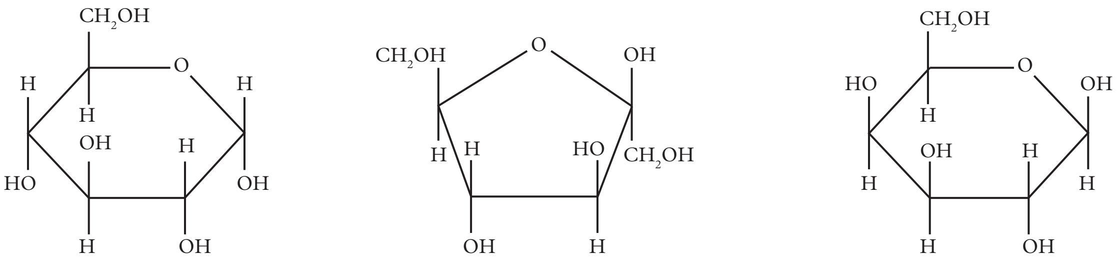
Glucose can be oxidised by specialised enzymes in plants or animals to produce carbon dioxide and water and release energy in the form of ATP. This process is aerobic respiration, which follows the following equation:

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

$$\text{glucose} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

The equation for aerobic respiration of glucose is the same as the equation for the complete combustion of glucose at standard laboratory conditions.

Plants can also convert glucose into other **monosaccharides** such as fructose and galactose. Glucose, fructose and galactose are structural isomers of each other – they all have the same molecular formula ($C_6H_{12}O_6$) but different structural formulas. All three monosaccharides are found in plants.



Glucose **Fructose** **Galactose**

FIGURE 1 Glucose, fructose and galactose are all found in plants.

Long-term energy stores

Glucose can be stored long-term in polymers; for example, as **starch** in plants and as **glycogen** in animals. Both starch and glycogen act as energy reserves that can be broken down (by water, in a hydrolysis reaction) to produce glucose monomers. You will learn about hydrolysis reactions in Chapter 9. Under normal conditions, neither starch nor glycogen are particularly soluble in water.

starch
the stored form of glucose in plants

glycogen
the stored form of glucose in animals

Glycogen

Animals store excess glucose as glycogen in the muscles and liver. Glycogen is a highly branched polymer made from glucose monomers. When the presence of glucose is low, glycogen can be converted back into glucose monomers.

Starch

Plants store excess glucose as starch, which is a polymer of glucose. Like glycogen in animals, starches are also formed by polymerisation reactions of glucose monomers and can be broken back down into glucose monomers.

There are two types of starch: **amylose** and **amylopectin**. Both are found in plants, but only amylopectin is digestible by humans due to its highly branched structure. Humans digest amylose much more slowly than amylopectin. This is because the straight structure of amylose makes it more difficult for the amylase enzyme to break the bonds holding the molecule together.

Compare how different starches increase blood glucose levels to different extents in Challenge 2.4.

amylose
a type of plant starch that is linear in structure

amylopectin
a type of plant starch that is branched in structure

2.4 Challenge
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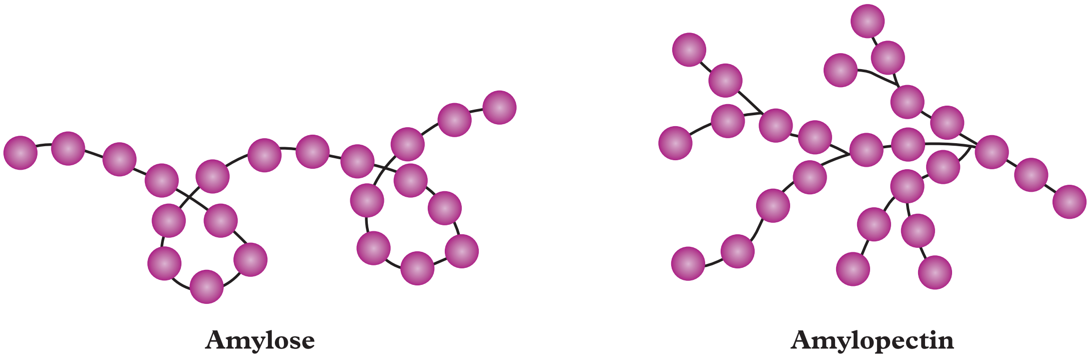

FIGURE 2 The two forms of starch: amylose and amylopectin

Study tip

You do not need to remember the chemical structures of glucose, glycogen and starch – they are provided to you in the data book.

TABLE 1 Comparison of the properties of the starches: amylose and amylopectin

Property	Amylose	Amylopectin
Structure	Straight-chain polymer of α-glucose units	Branched-chain polymer of α-glucose units
Digestion	Digested slowly by humans because the starch chains are straight – there is very little branching	Digested more readily by humans because the starch chains are more branched
Solubility in water	Less soluble than amylopectin	More soluble than amylose
Contribution to total starch in plants	~25% (but this varies)	~75% (but this varies)

2.4 REAL-WORLD CHEMISTRY

Amylomaize, a type of high-amylose corn

Amylomaize is a variety of corn plant that contains a gene that gives it an unusually high proportion of amylose starch compared to amylopectin starch. While other corn varieties have around 25% amylose starch and 75% amylopectin starch, amylomaize contains 50–94% amylose. This changes both the texture and the nutritional properties of the corn.

The amylomaize starch can also be made into an edible bioplastic film for wrapping food items. Astronauts on the Apollo missions in the late 1960s and early 1970s took food wrapped in amylomaize bioplastic film to prevent crumbs from floating off and damaging spacecraft equipment.

FIGURE 3 Amylomaize

Apply your understanding

1 Suggest two properties of amylomaize bioplastic that make it suitable as an edible food packaging material.

2 Suggest one potential health benefit of eating amylomaize instead of traditional maize.

2.4 CHECK YOUR LEARNING

Describe and explain

1 Identify the type of energy produced from the aerobic cellular respiration reaction.

2 Write the balanced chemical equation for aerobic cellular respiration in human cells.

3 List three monosaccharides with the chemical formula $C_6H_{12}O_6$.

Apply, analyse and compare

4 Compare the structure of amylose and amylopectin.

5 Explain why humans digest amylose more slowly than amylopectin.

FIGURE 4 Foods like fresh pasta are rich in amylopectin.

2.5

Producing bioethanol

KEY IDEAS

In this topic, you will learn that:

- bioethanol is produced by the fermentation of glucose extracted from plants:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$$

- bioethanol must be distilled to a high purity before it can be used as a transportation fuel in combustion engines
- bioethanol is renewable and can help reduce our reliance on fossil fuels, but it also uses farmland and water that could otherwise be used for food production.

Anaerobic fermentation

In Topic 2.4, you learnt about the oxidation of glucose in aerobic cellular respiration. The other form of cellular respiration is called anaerobic fermentation, where glucose is broken down to produce energy in the absence of oxygen.

In plants and yeasts, glucose can be broken down to produce ethanol and carbon dioxide in the absence of oxygen. This reaction releases chemical energy, and is shown below:

$$\begin{array}{cccc} C_6H_{12}O_6(aq) \rightarrow & 2C_2H_5OH(l) & + & 2CO_2(g) \\ \text{1 glucose} \rightarrow & \text{2 ethanol} & + & \text{2 carbon dioxide} \end{array}$$

The production of ethanol and carbon dioxide from fermentation has been used by humans for thousands of years to make products including bread and beer. It is now also being used to produce ethanol naturally, which can be used as a fuel.

Study tip

Combustion of any fuel, including bioethanol, produces water vapour as a product. Water vapour is a greenhouse gas. Unlike the emission of carbon dioxide, the emission of water vapour is not a major concern because the water vapour quickly condenses to liquid when it rains.

Producing bioethanol

Bioethanol, $C_2H_5OH(l)$, is a renewable substitute for petrol derived from the fermentation of plants. Bioethanol has several advantages over petrol derived from fossil fuels:

- **Renewability** – the plants required to produce bioethanol can be replenished in a relatively short period of time.
- **Decreased dependence on fossil fuels** – bioethanol can be produced locally and in small batches, which reduces our dependence on fossil fuels.
- **Lower carbon dioxide gas emissions** – it is relatively carbon neutral. The plants required to produce bioethanol capture carbon dioxide gas from the atmosphere via photosynthesis as they grow. The overall carbon dioxide gas emissions resulting from the production and combustion of bioethanol (or the 'life cycle') can therefore be lower than those produced by using petrol from fossil fuel sources.

The production of bioethanol is summarised in the flowchart below.

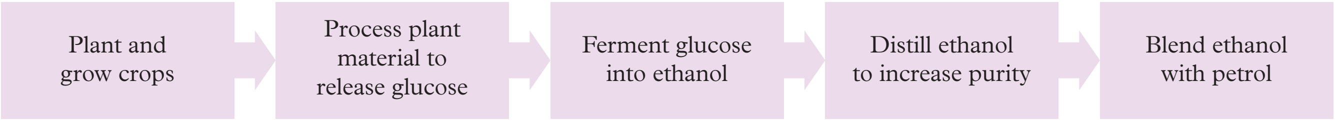

FIGURE 1 The production of bioethanol

Obtaining glucose for bioethanol production

To produce bioethanol, plants – the sources of glucose – need to be grown. Ideally, plants should grow quickly, with high photosynthetic efficiency, and produce large amounts of glucose with minimal resources.

In cold regions, sugar beet or corn are usually preferred. One example is the USA, where corn is grown for bioethanol production. In warmer regions, like Brazil, it is usually sugar cane. Of the three plants, sugar cane yields the most fermentable glucose, but it does not grow well in cold areas.

Mechanical processing releases some of the sugars from the plant into aqueous solution. However, much of the glucose is stored as polysaccharides, such as cellulose, hemicellulose and lignin, in the plant cell walls. These cell wall polysaccharides can be broken down into an aqueous mixture of fermentable glucose.

FIGURE 2 Biomass sources for bioethanol production: **a** sugar beet, **b** corn and **c** sugar cane

Fermenting glucose into ethanol

Once the mixture of glucose has been obtained, the glucose is fermented anaerobically (in the absence of oxygen) by yeast enzymes to produce ethanol and carbon dioxide gas.

This process is very similar to the fermentation reactions used in the production of bread and wine. In breadmaking, the desired product is carbon dioxide gas. This forms bubbles in the dough before it sets in the heat of the oven, giving the pleasant, spongy texture of leavened bread. The ethanol in breadmaking evaporates in the heat of the oven. In winemaking, the desired product of this same reaction is ethanol, the active ingredient in wine.

Distilling to high purity

Before ethanol is used as a fuel, it must be purified to remove almost all of the water from the ethanol mixture. This is done by distillation, which separates mixtures of liquids by their different boiling points. Water boils at 100°C but ethanol boils at just 78°C. By heating the mixture to slightly above 78°C, it is possible to evaporate the ethanol but not the water. The ethanol gas can then be condensed and collected in a separate container at a high purity. The purified ethanol produced from plants is called **bioethanol**.

bioethanol
ethanol produced from plants

Adding ethanol to petrol

Pure bioethanol damages petrol engines, so bioethanol is usually added to petrol at a low concentration – typically less than 10%. The resulting 90% petrol + 10% bioethanol mixture is sold in Australia as E10 fuel and is available at petrol stations. This can be added to a petrol engine without causing damage. Even though the bioethanol content of E10 fuel is only 10%, using E10 can help to reduce our reliance on fossil fuels slightly.

TABLE 1 Comparing bioethanol with petrol

Fuel	Advantages	Disadvantages	Approximate global production (2020)
Petrol	• Higher energy content • Cheap petrol can lower food prices, alleviating poverty	• Non-renewable • Requires fossil fuels • Releases carbon dioxide, a greenhouse gas	4.4 trillion litres/year
Bioethanol	• Renewable • Can be produced in places without access to fossil fuels	• Lower energy content • Bioethanol production can raise food prices by competing with agriculture for land and resources	100 billion litres/year

Practise writing balanced equations for combustion reactions in Skill drill 2.5, then apply your understanding of intermolecular forces to the properties of petrol and bioethanol in Real-world chemistry 2.5 and Challenge 2.5.

2.5 Skill drill
Find me in your obook pro

2.5 Real-world chemistry
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2.5 CHALLENGE

Why is ethanol soluble in both water and petrol?

Water is a polar substance that dissolves polar solutes such as sugar. Petrol is a non-polar substance that dissolves non-polar solutes such as hexane. Surprisingly, ethanol dissolves in water and in petrol. In terms of the intermolecular bonds, answer the following questions.

1 Explain how ethanol can dissolve in water. Include a labelled diagram in your answer.
2 Explain how ethanol can dissolve in petrol.
3 Explain whether you would expect the solubility of ethanol to be higher in petrol or in water.

2.5 CHECK YOUR LEARNING

Describe and explain

1 Explain what E10 fuel is.
2 Explain what determines whether a region will grow sugar cane or corn for bioethanol production.
3 Write a balanced chemical equation, including states, for the fermentation of glucose into ethanol and carbon dioxide.

Apply, analyse and compare

4 A number of the carbohydrates in plants are present as polymers in the cell walls. List some of the other polysaccharides.
5 Suggest how the polysaccharides in Question 4 can be broken down into glucose for fermentation into ethanol.

Design and discuss

6 Design a distillation process that would allow the purification of bioethanol.

2.6 Comparing exothermic and endothermic reactions

KEY IDEAS

In this topic, you will learn that:

- energy is required to break chemical bonds and energy is released when forming new chemical bonds
- endothermic reactions absorb heat energy from the environment, reducing temperature; and exothermic reactions release heat energy into the environment, increasing temperature
- enthalpy changes following the combustion of pure substances can be measured as ΔH in $kJ\,mol^{-1}$, and enthalpy changes of mixtures can be measured in $kJ\,g^{-1}$.

chemical energy
energy stored within chemical bonds

Chemical reactions involve the breaking and forming of chemical bonds. Energy stored within the chemical bonds of substances is called **chemical energy**. Energy is absorbed when bonds are broken and released when bonds are formed. The net amount of energy required and released during the chemical reaction can be either positive or negative. When fuels combust, they release energy (usually in the form of heat) that can be harnessed and used for different purposes.

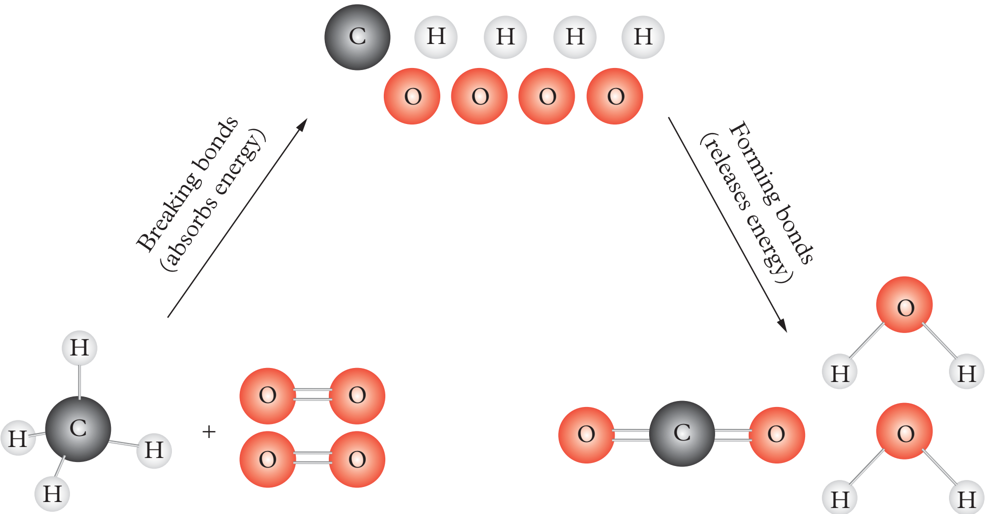


FIGURE 1 The relationship between energy and the breaking and forming of bonds – this reaction is the combustion of methane.

Endothermic and exothermic reactions

system
the chemical bonds holding together each reactant and product molecule

surroundings
the environment in which a chemical reaction is occurring

In a chemical reaction, an energy transfer occurs between the **system** (the bonds being broken and formed in the chemical reaction) and the **surroundings** (the environment in which the reaction takes place). For example, in the fermentation reaction of glucose:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$$

- the system is the covalent chemical bonds in the reactants that are broken and the covalent chemical bonds in the products that are formed
- the surroundings are the aqueous environment in which the reaction takes place.

Chemical reactions can be categorised into two groups based on whether the total chemical energy of products is less than or more than the total chemical energy of reactants. These two categories, **exothermic** and **endothermic** reactions, are used to describe whether energy from the system is being absorbed or released.

exothermic
a chemical reaction that is accompanied by the emission of heat energy to the surroundings

endothermic
a chemical reaction that is accompanied by the absorption of heat energy from the surroundings

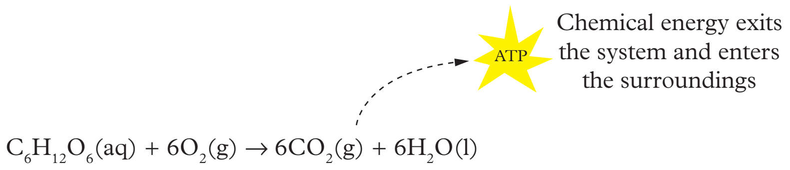
Exothermic reactions

If the chemical energy of products is less than the total chemical energy of reactants, any excess energy from the system will be released into the surroundings. In this type of reaction, energy has exited the system, so the reaction is exothermic (*exo* means 'outside' and *thermic* means 'temperature'). The net flow of energy from the system to the surroundings causes the surroundings to warm up. The system loses chemical energy and decreases in temperature.

Exothermic reactions follow the general formula:

$$\text{reactants} \rightarrow \text{products} + \text{energy}$$

For example, anaerobic cellular respiration is an exothermic reaction because excess chemical energy from the system is released into the surroundings as ATP. In these reactions, the products are more stable than the reactants.



$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

FIGURE 2 Aerobic cellular respiration is an exothermic reaction.

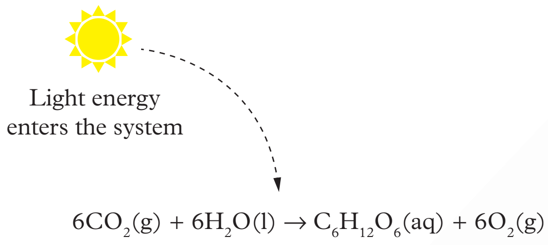
Endothermic reactions

If the chemical energy of products is more than the total energy chemical of reactants, energy from the surroundings will be absorbed into the system. In this type of reaction, energy has entered the system, so the reaction is categorised as endothermic (*endo* means 'inside'). The net flow of energy from the surroundings to the system causes the surroundings to cool down. The system gains chemical energy and increases in temperature.

Endothermic reactions follow the general formula:

$$\text{reactants} + \text{energy} \rightarrow \text{products}$$

For example, photosynthesis is an endothermic reaction because light energy from the surroundings enters the system. In these reactions, the products are not as stable as the reactants.



$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

FIGURE 3 Photosynthesis is an endothermic reaction.

FIGURE 4 Eucalyptus leaves use energy from the Sun for photosynthesis.

TABLE 1 Summary of exothermic and endothermic reactions

Feature	Exothermic reaction	Endothermic reaction
Definition	A chemical reaction in which heat energy is released	A chemical reaction in which heat energy is absorbed
Energy change	Energy from the system enters the surroundings	Energy from the surroundings enters the system
Enthalpy change (ΔH)	Negative (–); $\Delta H < 0$	Positive (+); $\Delta H > 0$
Temperature change	Increases	Decreases
Stability of products vs reactants	Products are more stable than the reactants	Products are less stable than the reactants
Example chemical reactions	Combustion, heat pack, cellular respiration	Photosynthesis, ice pack

Measuring changes in chemical energy (enthalpy)

enthalpy
a measure of chemical energy; represented by *H*

Enthalpy is the amount of energy stored within chemical bonds and is represented by the symbol H. Change in enthalpy is represented by ΔH, where the Δ symbol represents change and H is enthalpy or heat. The change in enthalpy from a chemical reaction can be either positive or negative. It is calculated using the following equation:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

As endothermic reactions absorb energy from the surroundings, the change in enthalpy is positive. As exothermic reactions release energy into the surroundings, the change in enthalpy is negative.

Representing changes using energy profile diagrams

energy profile diagram
a graph that shows the change in enthalpy as a reaction progresses

Energy changes in reactions, such as the change in enthalpy, can be plotted on **energy profile diagrams**. These diagrams show the difference between the amount of energy within the bonds of reactants and the amount of energy within the bonds of products.

In energy profile diagrams:

- the progression of the reaction is given on the x-axis and energy is given on the y-axis
- the plotted line represents how energy within bonds has changed over the course of a reaction
- reactants are found to the left side of the curve or 'peak' of the line; the height at which reactants sit is the energy required to break reactant bonds
- products are found to the right side of the curve or 'peak' of the line; the height at which products sit is the energy required to form product bonds
- the peak of the curve is the **transition state**, the period of the reaction where highly unstable, loosely bonded atoms react spontaneously to form the products
- substances at the bottom of the energy profile diagram have less chemical energy and stronger bonds
- substances at the top of the energy profile diagram have more chemical energy and weaker bonds.

transition state
an intermediate state in which atoms from the reactant molecules are loosely bonded together, unstable, and can spontaneously react to form new products

The shape of the energy profile diagram tells you whether a reaction is endothermic or exothermic. If reactants have more energy than products, energy is released to the surroundings and the reaction is exothermic. If reactants have less energy than products, energy is absorbed from the surroundings and the reaction is endothermic.

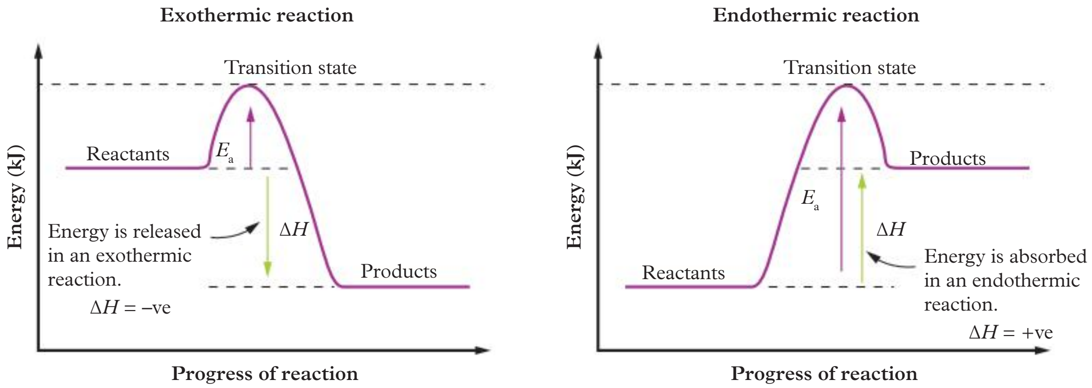


FIGURE 5 Energy profile diagrams for exothermic and endothermic reactions

Two important measurements can be made from an energy profile diagram:

1 ΔH is the change in enthalpy, measured in kJ mol^{-1}. It is the net energy absorbed (if $\Delta H > 0$) or released (if $\Delta H < 0$) during the reaction. ΔH is the vertical distance on the graph between reactants and products.

2 E_a is the **activation energy** measured in kJ mol^{-1}. It is the energy required to break the bonds in the reactants to initiate the chemical reaction. It is the vertical distance between the reactants on the graph and the transition state.

See Worked example 2.6 to learn how to draw an energy profile diagram.

Study tip

Heat energy is required to break chemical bonds. New chemical bonds then form spontaneously, releasing heat energy as they do so.

2.6 Worked example
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2.6 Worked example
Video demonstration

activation energy the minimum energy required to break the bonds between atoms in reactants

FIGURE 6 **a** The burning of firewood is an exothermic reaction. **b** The cooling of an ice pack is an endothermic reaction.

Changing states

Particles in gases are bonded more loosely and have a higher enthalpy than in liquids. The same is true when we compare particles in liquids with those in solids. This means that melting and evaporation are endothermic processes. The evaporation of sweat decreases the temperature of our skin because evaporation is an endothermic process.

- Endothermic state changes: **melting**, **vaporisation** and **sublimation**
- Exothermic state changes: **condensation**, **freezing** and **deposition**

The states of the reactants and products will alter the shape of the energy profile diagram (and also the ΔH) very slightly.

melting
the transition of a substance from solid to liquid state

vaporisation
the transition of a substance from liquid to gaseous state

sublimation
the transition of a substance directly from solid to gaseous state

condensation
the transition of a substance from gaseous to liquid state

freezing
the transition of a substance from liquid to solid state

deposition
the transition of a substance directly from gaseous to solid state

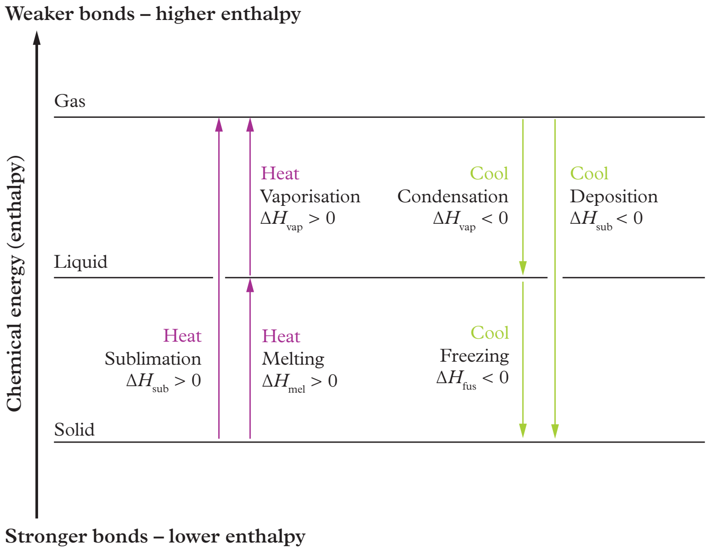

FIGURE 7 Enthalpy of solids, liquids and gases, and how enthalpy changes as they change states

Figure 8 shows how the different states of the H_2O products results in two different ΔH values for the following two chemical reactions:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) \qquad \Delta H = -802\,kJ\,mol^{-1}$$

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = -890\,kJ\,mol^{-1}$$

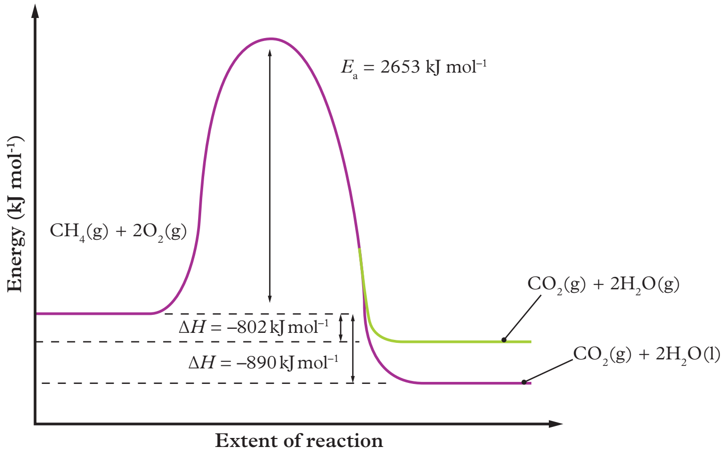

FIGURE 8 Different states of products result in different ΔH values in a chemical reaction.

Study tip

Be careful with the units for ΔH. Sometimes, change in enthalpy is expressed as $kJ\,mol^{-1}$, $kJ\,g^{-1}$, or even kJ alone. ΔH in kJ is used to show the total change in enthalpy, regardless of stoichiometry.

Experimental determination of enthalpy changes

The ΔH of the combustion of fuels can be calculated in two ways depending on whether the substances are pure or impure.

For pure substances, when the amount, in mol, is known:

$$\Delta H = \frac{q}{n}$$

For impure substances (mixtures):

$$\Delta H = \frac{q}{m}$$

where ΔH is the change in enthalpy in kilojoules per mole, q is the energy in kilojoules, n is the amount in mole and m is the mass in grams. These formulas are explored further in Topic 2.8 and Chapter 3. Try Challenge 2.6 to check your understanding of ΔH.

2.6 Challenge
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2.6 CHECK YOUR LEARNING

Describe and explain

1 Categorise the following state changes as endothermic or exothermic: vaporisation, melting, boiling, condensation, freezing, sublimation, deposition.

2 In an endothermic chemical reaction, identify whether the following changes are positive or negative.

  a Enthalpy

  b Temperature

3 Fill the blanks to complete the following sentence: Energy is ____________ when chemical bonds are broken, and energy is ____________ when chemical bonds are formed.

Apply, analyse and compare

4 Compare the enthalpy of the reactants and products in an endothermic reaction with the enthalpy of reactants and products in an exothermic reaction.

5 Given what we have learned about enthalpy in this chapter, predict whether stronger chemical bonds exist in the reactants or in the products of the equation for the complete combustion of octane.

6 Explain which of the following chemical reactions has a higher ΔH value.

- $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
- $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Design and discuss

7 Draw a fully labelled energy profile diagram for photosynthesis. Label the reactants, products, ΔH, E_a and transition state.

8 The chemical equations for aerobic cellular respiration and photosynthesis are the reverse of each other. Write an equation to define E_a (respiration) in terms of E_a (photosynthesis) and ΔH (photosynthesis).

FIGURE 9 The melting of ice is endothermic.

2.7

Limiting reactants and reagents

KEY IDEAS

In this topic, you will learn that:

- in any chemical reaction, one reactant may be limiting and the other reactants may be in excess.

reactant
a chemical species that undergoes change and is consumed during a reaction

reagent
a chemical species that participates in a reaction, but is not necessarily consumed

In a chemical reaction, **reactants** are the substances that initiate the reaction and are also consumed once the reaction is over. This is different to a **reagent**, a substance that is used to facilitate, detect or analyse a reaction, without necessarily being consumed by it.

For example, aerobic cellular respiration is catalysed by enzymes. Glucose and oxygen are reactants because they are consumed by the reaction, while the enzymes involved are reagents that facilitate the reaction but are not consumed.

Limiting and excess reactants and reagents

When a chemical reaction occurs, the reactants and/or reagents are not always present in the exact molar ratios required for the reaction. Usually, one substance is available in an insufficient amount and is said to be 'limiting'. The reaction will stop when this limiting substance runs out. Any chemical reaction (except for equilibrium reactions, which are covered in Chapter 6) can have **limiting reactants** and **limiting reagents**.

limiting reactant or reagent
the reactant or reagent that is completely consumed in a reaction

excess reactant or reagent
all other (non-limiting) reactants in a reaction; some amount of each excess reactant or reagent will remain unreacted at the end of the reaction

The limiting reactant or reagent can be determined by the molar ratios of the reaction. You should remember this from Units 1 and 2. Once the limiting substance has been identified, other reactants involved are then said to be 'in excess'. A quantity of these **excess reactants** and **excess reagents** will be left over or 'unreacted' by the end of the reaction (Figure 1).

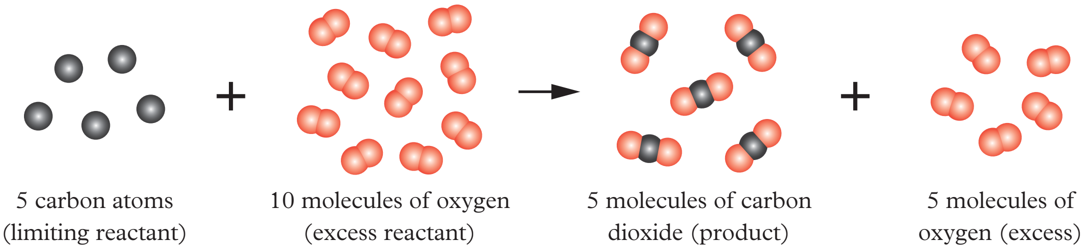

FIGURE 1 In the reaction above, the five carbon molecules limit the reaction. Since not all the oxygen molecules present can react to form carbon dioxide, oxygen is in excess.

For example, a reaction vessel could contain 1.0 mol of $CH_4(g)$ and 3.0 mol of $O_2(g)$. These gases react with each other in a 1 : 2 ratio according to the following equation:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

This means that the reaction can only proceed until 1.0 mol of $CH_4(g)$ depletes. The reaction will then finish, with 1.0 mol of $O_2(g)$ unreacted because there is no $CH_4(g)$ remaining to react with it.

Study tip

The terms *reactant* and *reagent* are often used interchangeably. But according to the dictionary, *reagents* include catalysts as well as reactants.

Determining limiting and excess reactants or reagents

To calculate the amount of product formed in a reaction, the molar ratio of the limiting substance must be used. This can be determined by examining the coefficients of reactants in a balanced equation. Consider the reaction between hydrogen and oxygen to form water:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Two moles of H_2 react with one mole of O_2 to produce two moles of H_2O. The molar ratio for this reaction is 2 H_2 : 1 O_2 : 2 H_2O. Because O_2 is the reactant with the lowest coefficient number, it is the limiting reactant. This also means that H_2 is the excess reactant.

Molar ratio can also be used to determine which reactant or reagent is limiting or in excess when the masses of reactants and reagents are given. If you recall from Year 11, $n = \frac{m}{M}$, where n is the amount in moles, m is mass and M is molar mass. If you divide n of a substance by the **stoichiometric ratio** in which it occurs in the reaction, you can then easily compare which substance is in excess and which substance is limiting. The limiting reactant will have a smaller $n \div$ ratio. This is shown in Worked example 2.7A.

stoichiometric ratio
the molar ratio in which the reactants react with each other (the coefficients in the balanced chemical equation)

2.7A WORKED EXAMPLE

DETERMINING THE LIMITING REACTANT

19.5 g of magnesium ribbon and 41.2 g of oxygen gas are reacted together in a vessel.

a Identify the limiting reactant.

b Calculate the mass of magnesium oxide formed.

Solution

<table>
<tr><th>Think</th><th>Do</th></tr>
<tr><td>Step 1: Write a balanced chemical equation for the reaction.</td><td>$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$</td></tr>
<tr><td>Step 2: Make an $m \mid M \mid n \mid$ molar ratio table and write the amounts given from the question into the table. Include the two masses that are given in the question, and the molar ratios based on the balanced chemical equation from Step 1.</td><td>
<table>
<tr><th></th><th>Mg</th><th>O_2</th><th>MgO</th></tr>
<tr><td>m</td><td>19.5</td><td>41.1</td><td></td></tr>
<tr><td>M</td><td>24.3</td><td>32.0</td><td></td></tr>
<tr><td>n (m/M)</td><td>0.80246</td><td>1.2844</td><td></td></tr>
<tr><td>Ratio</td><td>2</td><td>1</td><td>2</td></tr>
</table>
</td></tr>
<tr><td>Step 3: Divide the amount, in mol, of the reactants by the ratio in which they occur in the reaction ($n \div$ ratio) to determine which reactant is limiting and which is in excess.</td><td>$n(Mg) \div \text{ratio} = 0.40\,\text{mol}$
$n(O_2) \div \text{ratio} = 1.28\,\text{mol}$
$0.40\,\text{mol} < 1.28\,\text{mol}$
a Mg has the lowest $n \div$ ratio, so it is the limiting reactant and O_2 is in excess.</td></tr>
<tr><td>Step 4: Notice that Mg and MgO occur in the same ratio. This means the value of n you calculated for Mg (the limiting reactant) is the same as the value of n for MgO. Calculate M for MgO and use $m = n \times M$ to solve for the mass of MgO formed.</td><td>
<table>
<tr><th></th><th>Mg</th><th>O_2</th><th>MgO</th></tr>
<tr><td>m</td><td>19.5</td><td>41.1</td><td>32.3</td></tr>
<tr><td>M</td><td>24.3</td><td>32.0</td><td>40.3</td></tr>
<tr><td>n (m/M)</td><td>0.80246</td><td>1.2844</td><td>0.80246</td></tr>
<tr><td>Ratio</td><td>2</td><td>1</td><td>2</td></tr>
</table>
b $m(MgO) = 0.80246 \times 40.3$
$= 32.3\,\text{g}$ (3 sig fig)</td></tr>
</table>

You can also calculate the amount, in mol, of an excess reactant left over at the end of a reaction using the formula:

$$n(\text{excess reactant left over}) = \left(\frac{n_E}{\text{ratio}_E} - \frac{n_L}{\text{ratio}_L}\right) \times \text{ratio}_E$$

where E is the excess reactant and L is the limiting reactant. This is shown in Worked example 2.7B. Have a go at these calculations in Real-world chemistry 2.7.

2.7B WORKED EXAMPLE

CALCULATING THE QUANTITY OF EXCESS REACTANT LEFT AFTER A REACTION

Zinc metal reacts with hydrochloric acid according to the following equation:

$$\text{Zn(s)} + 2\text{HCl(aq)} \rightarrow \text{ZnCl}_2\text{(aq)} + \text{H}_2\text{(g)}$$

40.0 mL of hydrochloric acid with concentration 0.400 M is poured onto 1.10 g of zinc.

a Identify the limiting reactant.

b Calculate the amount, in mol, of hydrogen gas formed.

c Calculate the amount of excess reactant left over at the end of the reaction.

Solution

<table>
<tr><th>Think</th><th colspan="2">Do</th></tr>
<tr><td>Step 1: Calculate the amounts of hydrochloric acid and zinc, in mol.</td><td>$n(\text{HCl}) = c \times V$
$= 0.400 \times (40.0 \div 1000)$
$= 0.0160$ mol</td><td>$n(\text{Zn}) = \frac{m}{M}$
$= \frac{1.10}{65.4}$
$= 0.016\,820$ mol</td></tr>
<tr><td>Step 2: Make an m | n | ratio table and write the amounts given into the table. Include n(HCl) and n(Zn) calculated above, and the stoichiometric ratios provided in the question. Then, use n ÷ ratio to determine the limiting reactant.</td><td colspan="2"><table>
<tr><th></th><th>Zn</th><th>HCl</th><th>ZnCl_2</th><th>H_2</th></tr>
<tr><td>m</td><td>1.10</td><td></td><td></td><td></td></tr>
<tr><td>n</td><td>0.016 820</td><td>0.0160</td><td></td><td></td></tr>
<tr><td>ratio</td><td>1</td><td>2</td><td>1</td><td>1</td></tr>
<tr><td>n ÷ ratio</td><td>0.016 820</td><td>0.008 00</td><td></td><td></td></tr>
</table>
0.016 mol > 0.008 mol
a HCl has the lowest n ÷ ratio, so it is the limiting reactant.</td></tr>
<tr><td>Step 3: Use the limiting reactant data and the mol ratio to calculate the amount of H_2, in mol, and then the mass of H_2 formed.</td><td colspan="2">**b** $n(\text{H}_2) = \frac{1}{2} \times n(\text{HCl})$
$= \frac{1}{2} \times 0.0160$
$= 0.008$ mol
$m(\text{H}_2) = n \times M$
$= 0.008 \times (2 \times 1.0)$
$= 0.0160$ g (3 sig fig)
<table>
<tr><th></th><th>Zn</th><th>HCl</th><th>ZnCl_2</th><th>H_2</th></tr>
<tr><td>m</td><td>1.10</td><td></td><td></td><td>0.016</td></tr>
<tr><td>n</td><td>0.016 820</td><td>0.0160</td><td></td><td>0.008</td></tr>
<tr><td>ratio</td><td>1</td><td>2</td><td>1</td><td>1</td></tr>
<tr><td>n ÷ ratio</td><td>0.016 820</td><td>0.008 00</td><td></td><td></td></tr>
</table></td></tr>
<tr><td>Step 4: Use the formula $\left(\frac{n_E}{\text{ratio}_E} - \frac{n_L}{\text{ratio}_L}\right) \times \text{ratio}_E$ to calculate the quantity of excess reactant left at the end of the reaction.</td><td colspan="2">**c** $n(\text{excess Zn}) = \left(\frac{n_E}{\text{ratio}_E} - \frac{n_L}{\text{ratio}_L}\right) \times \text{ratio}_E$
$= \left(\frac{0.016\,820}{1} - \frac{0.0160}{2}\right) \times 1$
$= 0.008\,82$ mol (3 sig fig)</td></tr>
</table>

2.7 REAL-WORLD CHEMISTRY

Calculating exact amounts for a thermite reaction

Thermite is a mixture of a reactive metal powder with the oxide of a less reactive metal. Typically, a mixture of aluminium powder and iron(III) oxide powder are used. When ignited, a highly exothermic redox reaction occurs.

FIGURE 2 The enormous amount of heat generated when thermite is ignited is used to melt lengths of track together when building a railway.

Apply your understanding

1 Write a balanced chemical equation for the chemical reaction between aluminium and iron(III) oxide to produce iron and aluminium oxide.

2 Calculate the mass of aluminium powder required to react completely with 2.7 kg of iron(III) oxide.

2.7 CHECK YOUR LEARNING

Describe and explain

1 Define the term *limiting reactant*.

Apply, analyse and compare

2 Determine the limiting reactant when 2 mol of substance A is allowed to react with 3 mol of substance B in the equation 3A + 2B → 3C + 2D.

3 Calculate the amount, in mol, of excess reactant left over at the end of the reaction in Question 2 above.

4 4.7 g of magnesium ribbon is allowed to react with 6.9 g of oxygen gas in a container.
  - **a** Write a balanced equation for this reaction.
  - **b** Determine the limiting reactant.
  - **c** Calculate the amount, in mol, of excess reactant remaining at the end of the reaction.

5 61 g of propane gas is allowed to react with 70 g of oxygen gas in a container.
  - **a** Write a balanced equation for this reaction.
  - **b** Determine the limiting reactant.
  - **c** Calculate the amount, in mol, of excess reactant remaining at the end of the reaction.

Design and discuss

6 Calculate the volume of oxygen (at SLC) required to react completely with 98 mL of carbon monoxide in the following reaction:

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

2.8 Combustion of fuels

KEY IDEAS

In this topic, you will learn that:

- fuels can combust completely or incompletely, producing different products, different Δ*H* values and different flame characteristics.

complete combustion
the burning of a fuel in excess oxygen supply to produce carbon dioxide gas and water vapour

incomplete combustion
the burning of a fuel in limited oxygen supply to produce carbon monoxide gas, solid carbon (soot) and water vapour

Combustion reactions occur when fuels react with oxygen and release heat energy. This means all combustion reactions are exothermic. Fuels combust differently depending on the availability of oxygen. The amount of oxygen available determines whether the reaction will be **complete combustion** or **incomplete combustion**.

FIGURE 1 Combustion of gas producing a flame

Complete combustion

When oxygen is in excess, complete combustion occurs and all fuel molecules are oxidised. Complete combustion burns with a clean, blue flame and produces carbon dioxide gas and water vapour as products.

The complete combustion of hydrocarbon (such as octane) or alcohol fuels (such as ethanol) occurs according to the following equations:

hydrocarbon + oxygen → carbon dioxide + water
alcohol + oxygen → carbon dioxide + water

The complete combustion of carbon-based fuels (such as coal) occurs according to the following equation:

carbon + oxygen → carbon dioxide

Study tip

Make sure you remember the different products of incomplete and complete combustion.

Incomplete combustion

When oxygen is the limiting reactant, incomplete combustion will occur and not all fuel molecules will be completely reacted. In this situation, they will form other products, such as toxic carbon monoxide gas or solid carbon particles (soot). Incomplete combustion produces a yellow flame.

The incomplete combustion of hydrocarbon fuels, where you do not need to show the formation of carbon soot, occurs as follows:

hydrocarbon + oxygen → carbon monoxide + water
alcohol + oxygen → carbon monoxide + water

The incomplete combustion of carbon-based fuels occurs according to the following reaction equation:

carbon + oxygen → carbon monoxide

FIGURE 2 The incomplete combustion of hydrocarbons can produce carbon monoxide gas, which can be deadly.

2.8
PRACTICAL: CONTROLLED EXPERIMENT

What is the energy content of three liquid fuels? Go to your **Student obook pro**.

Flame characteristics

Bunsen burners have a hole at the base that can be opened or closed to control the availability of oxygen to the flame. When the hole is open, air can flow into the Bunsen burner and mix with the natural gas in plentiful supply to ensure complete combustion. This produces a blue flame and a high temperature. Closing the hole restricts the availability of oxygen in the Bunsen burner, resulting in incomplete combustion. This gives a characteristic yellow flame and produces tiny particles of soot.

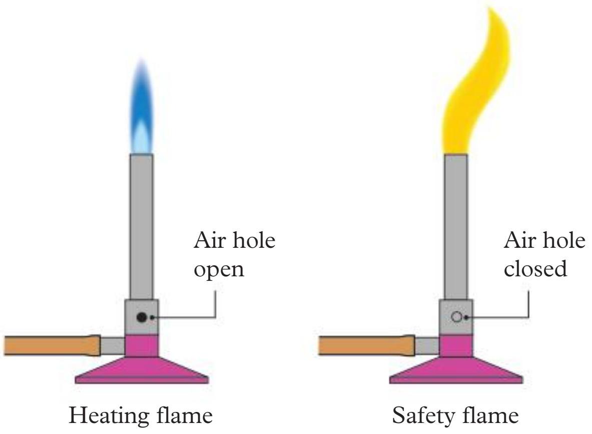

FIGURE 3 Comparing Bunsen burner flame colours

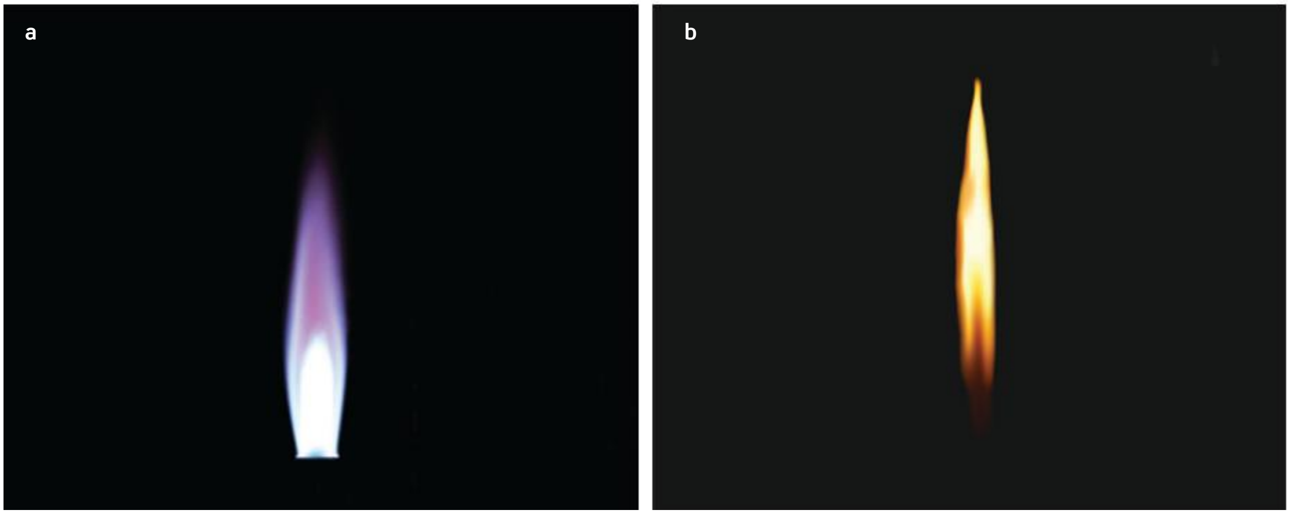

FIGURE 4 **a** Complete combustion produces a blue flame. **b** Incomplete combustion produces a yellow flame.

Heat of combustion

Heat of combustion is the energy released when a specific amount of a fuel is burnt during a complete combustion reaction. It is usually measured under standard laboratory conditions (SLC, 25°C and 100 kPa). This means that any water produced by combustion will be in liquid state.

In Topic 2.2, you learnt that the heat of combustion of common macromolecules is measured in $\mathrm{kJ\,g^{-1}}$. This is because these substances are usually complex mixtures of chemicals and not always pure. If a substance is in pure form, heat of combustion can be measured in $\mathrm{kJ\,mol^{-1}}$.

Table 1 on the next page shows the heats of combustion of common organic substances. As you learnt in Topic 2.6, the enthalpy of a reaction will change as the reaction progresses. The **enthalpy of combustion** is the change in reaction energy that occurs due to complete combustion. Enthalpy of combustion is a negative value due to the exothermic nature of combustion. You will learn more about how to use these values in calculations in the next chapter.

heat of combustion
the amount of heat energy released when a specific quantity of a fuel combusts

enthalpy of combustion
the change in enthalpy when a quantity of fuel undergoes complete combustion

TABLE 1 Common organic fuels and their combustion values

Organic substance	Formula	State at SLC	Molar heat of combustion ($kJ\,mol^{-1}$)	Enthalpy of combustion ($kJ\,mol^{-1}$)
Hydrogen	H_2	gas	282	–282
Methane	CH_4	gas	890	–890
Ethane	C_2H_6	gas	1560	–1560
Propane	C_3H_8	gas	2220	–2220
Butane	C_4H_{10}	gas	2880	–2880
Octane	C_8H_{18}	liquid	5460	–5460
Ethyne (acetylene)	C_2H_2	gas	1300	–1300
Methanol	CH_3OH	liquid	726	–726
Ethanol	C_2H_5OH	liquid	1360	–1360

Writing balanced incomplete and complete combustion reactions

The steps involved for balancing complete and incomplete combustion reactions are similar.

1 On the left side of the equation, write the reactants: the fuel with its correct state and oxygen gas, $O_2(g)$.

2 On the right side of the equation, write the products. Complete combustion has products $H_2O(l)$ and $CO_2(g)$. Incomplete combustion has products $H_2O(l)$ and $CO(g)$.

3 Balance the combustion equation by balancing the carbon atoms first, then the hydrogen atoms, and finally the oxygen atoms.

Examples of writing balanced complete and incomplete combustion equations are shown in Worked examples 2.8A and 2.8B, respectively.

2.8A Worked example
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2.8A Worked example
Video demonstration

2.8B Worked example
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2.8B Worked example
Video demonstration

Writing balanced thermochemical equations

Thermochemical equations are chemical equations that show a change in enthalpy. The change in enthalpy, ΔH, is often provided in the units $kJ\,mol^{-1}$. The value of ΔH will either be negative or positive depending on the type of reaction taking place. Endothermic reactions will have a positive ΔH value and exothermic reactions will have a negative ΔH value.

The enthalpy of combustion values given in Table 1 can be used to write thermochemical equations for complete combustion reactions. For example, Figure 5 shows a labelled thermochemical equation for the complete combustion of methane.

thermochemical equation
a chemical equation that shows the enthalpy change of the reaction

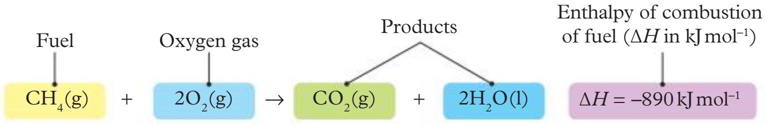


FIGURE 5 Thermochemical equation for the complete combustion of methane

The value shown represents one mole of the fuel (methane) reacting – note the $kJ\,mol^{-1}$ units. If the coefficients in the reaction change, so will the enthalpy (in $kJ\,mol^{-1}$) and total energy (in kJ). For example, if all the coefficients of this reaction were doubled, there would be two moles of CH_4 reacting, which means twice the amount of energy is released. However, in your calculations, you must use the molar ΔH, or: $q = \frac{\Delta H \times n}{\text{coefficient}}$.

Study tip

If the coefficient of the fuel in the balanced equation is greater than 1, the ΔH must be multiplied by the coefficient of the fuel.

If a thermochemical equation is reversed, the magnitude of ΔH will remain the same. However, previously negative ΔH values will be positive (as the reversed reaction is now endothermic) and previously positive ΔH values will be negative (as the reversed reaction is now exothermic).

Enthalpy change (ΔH in kJ mol^{-1}) is positive.

$$CO_2(g) + 2H_2O(l) \rightarrow CH_4(g) + 2O_2(g) \quad \boxed{\Delta H = +890\,\text{kJ mol}^{-1}}$$

FIGURE 6 Reversing a combustion reaction changes the enthalpy of combustion value from negative to positive and vice versa.

2.8C WORKED EXAMPLE

WRITING A BALANCED THERMOCHEMICAL EQUATION

Write the balanced thermochemical equation for the complete combustion of methanol. Then, determine the heat of combustion, in kJ, when 5.00 g of methanol is completely combusted.

Solution

Think	Do
Step 1: Write a balanced equation for the complete combustion of methanol.	$2CH_3OH(l) + 3O_2(g) \rightarrow 4H_2O(l) + 2CO_2(g)$
Step 2: Use Table 1 to determine the enthalpy of combustion for methanol. If the coefficient of the fuel in the balanced equation is greater than 1, the ΔH must be multiplied by the coefficient of the fuel. Put this together with your balanced equation. Remember that the combustion reaction is exothermic, so the ΔH is negative.	$2CH_3OH(l) + 3O_2(g) \rightarrow 4H_2O(l) + 2CO_2(g)$ $\Delta H = -726 \times 2 = -1452\,\text{kJ mol}^{-1}$
Step 3: Calculate the amount, in mol, of methanol combusted.	$n(CH_3OH) = \frac{m}{M}$ $= \frac{5.00}{12.0 + (4 \times 1.0) + 16.0}$ $= 0.156\,\text{mol}$
Step 4: Multiply the ΔH value with the amount of methanol combusted to determine the heat of combustion, in kJ.	$q = \frac{\Delta H \times n}{\text{coefficient}}$ $= \frac{1452 \times 0.156}{2}$ $= 113\,\text{kJ}$ (3 sig fig)

2.8 CHECK YOUR LEARNING

Describe and explain

1 List the possible products of incomplete combustion.

2 Describe how heat of combustion differs from enthalpy of combustion.

Apply, analyse and compare

3 Compare complete and incomplete combustion reactions.

4 Write balanced chemical equations for:
  a the complete combustion of octane
  b the incomplete combustion of butane
  c the complete combustion of glucose
  d the incomplete combustion of ethyne.

5 Write a balanced thermochemical equation for:
  a the complete combustion of ethanol
  b the reverse of the reaction in part **a**.

6 Explain what has happened to the enthalpy of combustion values between the two reactions in Question 5, and why this is so.

7 Calculate the change in enthalpy, in kJ, when:
  a 4.50 mol of propane is completely combusted
  b 14.2 g of butane is completely combusted.

Design and discuss

8 Discuss how incomplete combustion reactions could pose safety concerns in an industrial factory setting.

CHAPTER 2 Review

Chapter summary

2.1
- A fuel is a substance that can be reacted with another substance (e.g. combusted with oxygen in the air) to produce useful energy.
- Fossil fuels are non-renewable mixtures of hydrocarbons extracted from the ground, such as coal, oil and natural gas.
- Biofuels are renewable fuels made from plants that can be replaced by natural processes within a relatively short period of time.

2.2
- Carbohydrates, proteins and lipids (fats and oils) are fuel sources for the human body.
- The energy provided to the body by fuel sources is measured in kJ g^{-1}.

2.3
- Photosynthesis is the process that converts light energy into chemical energy.

2.4
- Glucose is the main carbohydrate used as an energy source by plants and animals.

2.5
- Bioethanol is produced by the fermentation of glucose extracted from plants.
- Bioethanol must be distilled to a high purity before it can be used as a transportation fuel in combustion engines.
- Bioethanol is renewable and can help reduce our reliance on fossil fuels, but it uses farmland and water that could otherwise be used for food production.

2.6
- Energy is required to break chemical bonds and is released when forming new chemical bonds.
- Endothermic reactions absorb heat energy from the environment, reducing temperature. Exothermic reactions release heat energy to the environment, increasing temperature.

2.7
- In any chemical reaction, at least one reactant will be limiting, and the other reactants will be in excess.

2.8
- Fuels can combust completely or incompletely, producing different products, different ΔH values and different flame characteristics.

Key formulas

Photosynthesis	$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$
Cellular respiration	$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$
Fermentation of glucose	$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$
Change in enthalpy	$\Delta H = H_{\text{products}} - H_{\text{reactants}}$
Change in enthalpy of pure substances	$\Delta H = \frac{q}{n}$
Change in enthalpy of impure substances	$\Delta H = \frac{q}{m}$
Amount of excess reactant	$n(\text{excess reactant}) = \left(\frac{n_E}{\text{ratio}_E} - \frac{n_L}{\text{ratio}_L}\right) \times \text{ratio}_E$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as ‘Confidently’, ‘Mostly’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Mostly	Not really	Revision link
• describe the difference between fossil fuels and biofuels with reference to their renewability	☐	☐	☐	Go back to Topic 2.1
• determine how much energy in $kJ\,g^{-1}$ the body can obtain by eating amounts of carbohydrates, lipids and proteins	☐	☐	☐	Go back to Topic 2.2
• write the photosynthesis equation and describe how photosynthesis converts light energy into chemical energy and provides oxygen and glucose needed for respiration	☐	☐	☐	Go back to Topic 2.3
• describe how glucose is oxidised and write the equation for cellular respiration	☐	☐	☐	Go back to Topic 2.4
• explain how bioethanol is produced by the fermentation of glucose	☐	☐	☐	Go back to Topic 2.5
• compare endothermic and exothermic reactions	☐	☐	☐	Go back to Topic 2.6
• determine the limiting and excess reactants and reagents in chemical equations	☐	☐	☐	Go back to Topic 2.7
• write balanced thermochemical equations for complete and incomplete combustion reactions	☐	☐	☐	Go back to Topic 2.8

Revision questions

Multiple choice

1 Which of the following fuels is not considered renewable?

A Nuclear energy
B Biodiesel
C Bioethanol
D Biogas

2 Approximately how long ago did Australia’s large oil deposits form?

A 10 thousand years ago
B 100 thousand years ago
C 10 million years ago
D 100 million years ago

3 Which of the following fuels has the greatest molar mass?

A Biodiesel
B Bioethanol
C Biogas
D Petrol

4 Which of the following provides the greatest amount of energy from digestion?

A 18.0 g of fat
B 24.0 g of protein
C 26.0 g of carbohydrate
D One boiled egg that provides 590 kJ of energy

5 If 100.0 g of cashew nuts contains 44.0 g fat, 30.0 g carbohydrate and 18.0 g protein, the energy obtained from digestion is closest to:

A 306 kJ.
B 480 kJ.
C 1628 kJ.
D 2414 kJ.

FIGURE 1
Cashew nuts

6 Proteins produce approximately 24 kJ g^{-1} during combustion but only 17 kJ g^{-1} during digestion because:

A enzymes in the body convert proteins into fats.

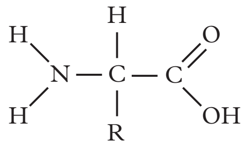
B digestion involves incomplete combustion.

C heat is lost to the surroundings.

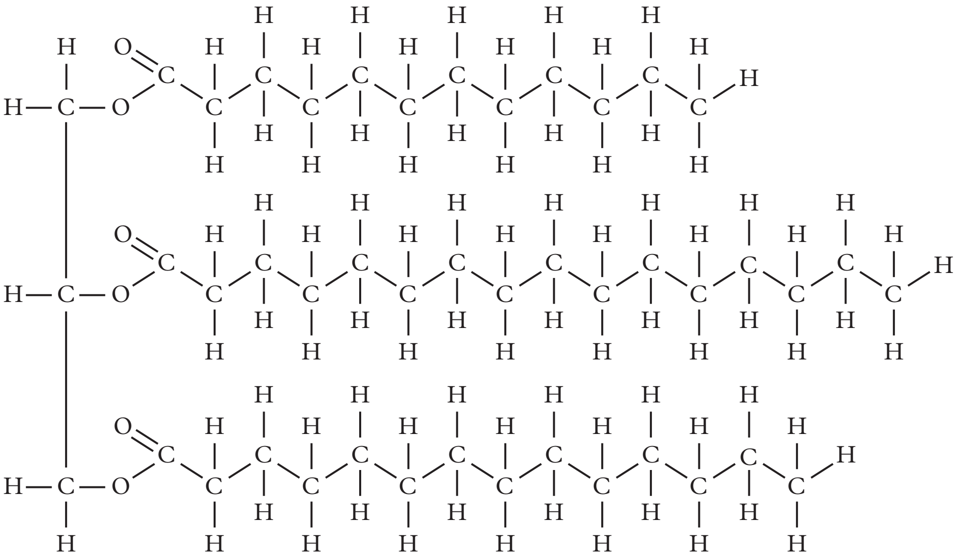
D the body's enzymatic processes are not able to oxidise and absorb all the nutrients completely.

7 Which of the following structural formulas shows an amino acid?

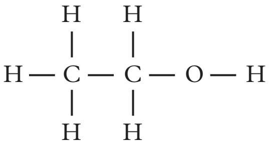
A



B

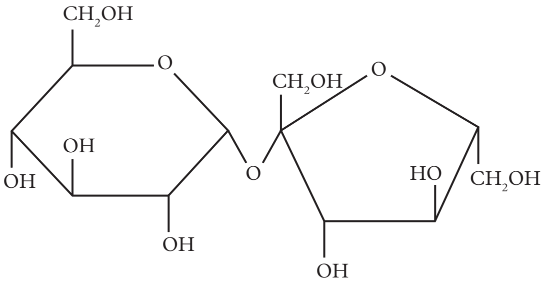


C



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & O-H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

D



8 Use the nutritional information below to calculate the total energy obtained by digestion when a person eats an entire pineapple of mass 905 g.

Nutrient	Value per 905 g pineapple (g)
Fat	1.1
Carbohydrate	119
Protein	4.9

A 5.72 kJ

B 2.03 kJ

C 2.24 kJ

D 2.03 MJ

FIGURE 2 Pineapples contain fats, carbohydrates and proteins.

9 Which of the following factors does not increase the rate of photosynthesis in plants?

- **A** Light intensity
- **B** Atmospheric carbon dioxide concentration
- **C** Temperature
- **D** Atmospheric nitrogen concentration

10 Pure bioethanol must not be used in petrol engines. Why is this?

- **A** It can damage the engine.
- **B** It is too expensive to use 100% bioethanol.
- **C** Its vapour pressure is too low.
- **D** The statement is incorrect. Pure bioethanol can be used in petrol engines.

Short answer

Describe and explain

11 Write a balanced chemical equation for cellular respiration of glucose ($C_6H_{12}O_6$) including states.

12 Write a balanced chemical equation for photosynthesis in the presence of sunlight.

13 Two different structures of starch are shown below.

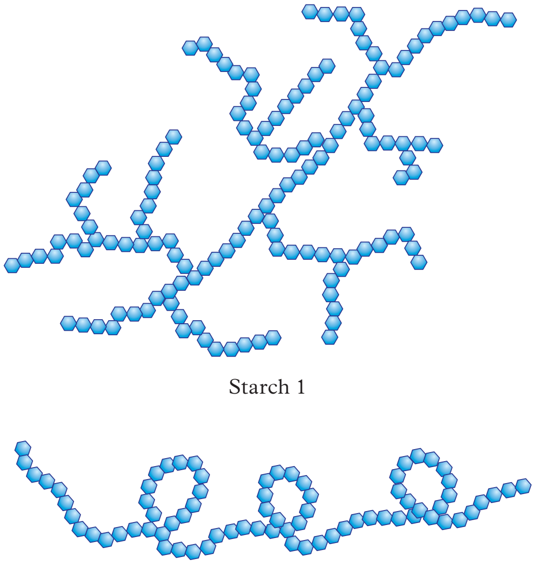


FIGURE 3 Starch chains

- **a** Suggest which image shows amylose starch.
- **b** Suggest which image shows amylopectin starch.
- **c** Suggest which starch is more soluble in water (and explain your choice).

14 Explain what a fuel is.

15 Explain why the oxidation of the same masses of carbohydrate, protein and lipid will release different amounts of energy.

FIGURE 4 Oxidation of different foods releases different amounts of energy.

16 Identify the green chemistry principle that is addressed by producing E10 fuel. Justify your choice.

17 Write a balanced chemical equation for the fermentation of ethanol by enzymes.

18 Describe the process by which E10 fuel is produced.

Apply, analyse and compare

19 Describe three advantages of bioethanol over petrodiesel.

20 Describe one disadvantage of bioethanol over petrodiesel.

21 The thermochemical equation for the complete combustion of solid carbon is shown below:

$$C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -394\,kJ\,mol^{-1}$$

- **a** Calculate the ΔH, in kJ, for the reaction of 4.0 mol C with 4.0 mol O_2.
- **b** Calculate the ΔH, in kJ, for the reaction of 10.0 mol C with 10.0 mol O_2.
- **c** Determine whether the above reactions are endothermic or exothermic.

22 Identify which of the following state changes are endothermic.

I Melting
II Boiling
III Condensation
IV Freezing
V Sublimation
VI Deposition

FIGURE 5 Condensation on a glass window

23 Bioethanol combusts in excess oxygen to produce water vapour and carbon dioxide gas.

- **a** Write a balanced chemical equation for this reaction.
- **b** Compare the enthalpy of reactants with the enthalpy of the products.
- **c** Calculate the heat energy released, in kJ, when 1.00 kg of bioethanol is completely combusted.

24 Propane gas, $C_3H_8(g)$, combusts in excess oxygen to produce water vapour and carbon dioxide gas.

- **a** Write a balanced chemical equation for this reaction.
- **b** The heat of combustion of propane is 50.5 $kJ\,g^{-1}$. Calculate the amount of energy, in kJ, released from the complete combustion of 50.0 g propane gas at SLC.

25 State the differences between a blue flame and a yellow flame in a Bunsen burner. In your answer, refer to the hole (open/closed), the temperature of the flame (higher/lower) and the chemical formulas of the products formed.

FIGURE 6 Combustion of propane produces heat energy used on a stove.

FIGURE 7 The blue and yellow Bunsen burner flames

26 The energy profile diagram for the combustion of nonane at SLC is shown below.

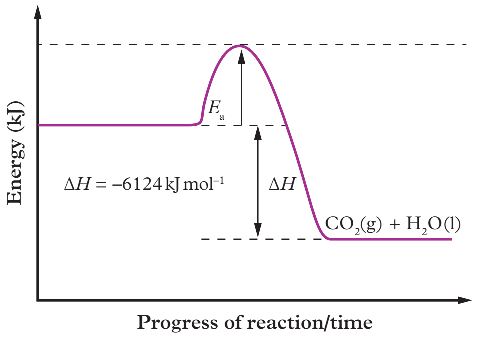


FIGURE 8 The combustion of nonane

a Write a balanced chemical equation for the combustion of nonane in excess oxygen.

b Identify whether the reactants or the products have the greater enthalpy.

c Identify whether the reactants or the products have weaker bonds.

27 4.0 mol A is allowed to react with 1.1 mol B in the equation: 3A + B → 3C + D.

a Determine the limiting reactant.

b Calculate the amount of excess reactant remaining when the reaction is complete.

28 0.040 mol X is allowed to react with 0.10 mol Y in the equation X + 2Y → C + D.

a Determine the limiting reactant.

b Calculate the amount of excess reactant remaining when the reaction is complete.

29 Ethene gas combusts with oxygen gas to produce water vapour and carbon dioxide gas.

a Write a balanced chemical equation for this reaction.

b Determine the limiting reactant if 14.0 g ethene gas and 44.0 g oxygen gas are allowed to react in a sealed container.

c Calculate the number of grams of excess reactant left over at the end of the reaction above.

30 Hydrochloric acid reacts with calcium carbonate according to the equation:

$$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Determine the limiting reactant if 0.050 mol HCl is added to 0.40 mol $CaCO_3$.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER 3

Measuring changes in chemical reactions

KEY KNOWLEDGE

- → calculations related to the application of stoichiometry to reactions involving the combustion of fuels, including mass–mass, mass–volume and volume–volume stoichiometry, to determine heat energy released, reactant and product amounts and net volume or mass of major greenhouse gases (CO_2, CH_4 and H_2O), limited to standard laboratory conditions (SLC)
- → the use of specific heat capacity of water to approximate the quantity of heat energy released during the combustion of a known mass of fuel and food
- → the principles of solution calorimetry, including determination of calibration factor and consideration of the effects of heat loss; analysis of temperature–time graphs obtained from solution calorimetry
- → energy from fuels and food:
  - calculation of energy transformation efficiency during combustion as a percentage of chemical energy converted to useful energy
  - comparison and calculations of energy values of foods containing carbohydrates, proteins and fats and oils

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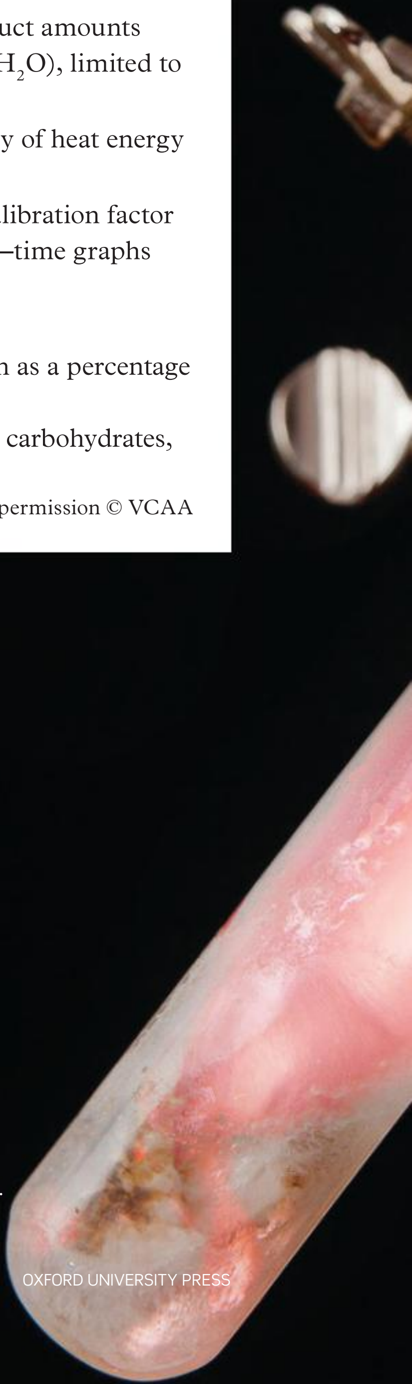
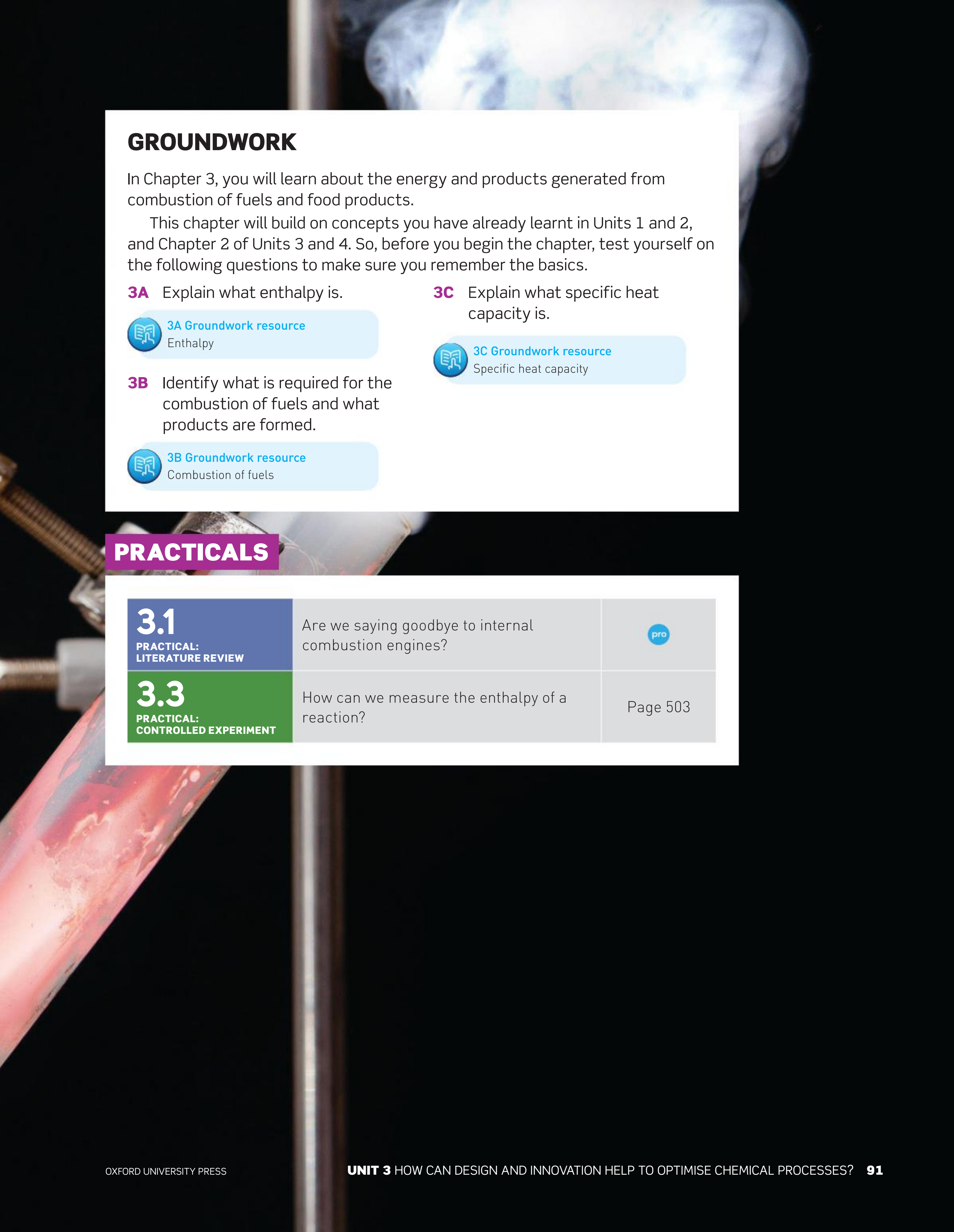


FIGURE 1 Energy is released from the combustion of a jelly baby. This is also called the screaming jelly baby reaction.



GROUNDWORK

In Chapter 3, you will learn about the energy and products generated from combustion of fuels and food products.

This chapter will build on concepts you have already learnt in Units 1 and 2, and Chapter 2 of Units 3 and 4. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

3A Explain what enthalpy is.

3A Groundwork resource
Enthalpy

3B Identify what is required for the combustion of fuels and what products are formed.

3B Groundwork resource
Combustion of fuels

3C Explain what specific heat capacity is.

3C Groundwork resource
Specific heat capacity

PRACTICALS

3.1 **PRACTICAL: LITERATURE REVIEW**	Are we saying goodbye to internal combustion engines?	pro
3.3 **PRACTICAL: CONTROLLED EXPERIMENT**	How can we measure the enthalpy of a reaction?	Page 503

3.1 Stoichiometry involving the combustion of fuels

KEY IDEAS

In this topic, you will learn that:

- stoichiometry can be used to determine the mass and volume of substances involved in combustion reactions
- the net volume or mass of major greenhouse gases (CO_2, CH_4 and H_2O) generated from the combustion of fuels can be calculated using stoichiometry
- the energy released in combustion reactions can be calculated using heat of combustion.

combustion
an exothermic chemical reaction with oxygen that produces heat, carbon dioxide and water

greenhouse gases
gases such as CO_2, H_2O and CH_4, which contribute to the greenhouse effect

Chemical reactions involve a change in energy to make new substances. In a chemical reaction, the bonds of reactants are broken, and the atoms rearrange to form new bonds in products.

One important type of chemical reaction is the **combustion** of fuels. Combustion of hydrocarbons is an exothermic reaction with oxygen that generates heat, carbon dioxide (CO_2) and water (H_2O). Two of these products (carbon dioxide and water) are **greenhouse gases**. You might remember from Units 1 and 2 that greenhouse gases are atmospheric gases that trap infrared radiation and prevent it from leaving the atmosphere.

The third major greenhouse gas is methane (CH_4), which is produced naturally and in high quantities by wetlands, but also by humans through agriculture. The current percentages of the major greenhouse gases in the atmosphere are:

- water: 0–4%
- carbon dioxide: 0.0420%
- methane: 0.000 17%.

FIGURE 1 Combustion of fuel in a car engine releases carbon dioxide, a greenhouse gas.

FIGURE 2 Human activities such as agriculture contribute significantly to atmospheric greenhouse gases.

3.1
PRACTICAL: LITERATURE REVIEW

Are we saying goodbye to internal combustion engines? Go to your **Student obook pro**.

Greenhouse gases can contribute significantly to global warming and climate change. Their amounts in the atmosphere can change over time due to human activities. During the 2020 COVID-19 lockdowns, global greenhouse gas emissions were reduced by 6.4%, or 2.3 billion tonnes. The biggest decrease in emissions came from international travel bans, which prevented many flights. This reduced the need to burn fossil fuels, which power aeroplane engines. Other sectors experienced no decrease in greenhouse gas emissions, as many people worked from home (requiring electricity, heating and water).

Study tip

Within your study design, the calculation of the net volume or mass of greenhouse gases is limited to **standard laboratory conditions (SLC)**: 25°C and 100 kPa. Because water is a liquid and not a gas at SLC, it should not be considered a greenhouse gas at SLC.

SLC
standard laboratory conditions of 25°C and 100 kPa

Solving stoichiometry problems

Fuels are expressed in quantities of mass, volume or density. When purchasing fuel to fill a car's fuel tank, you deal with units of volume (e.g. litres) since fuel is a liquid. Gases found in gas cylinders (e.g. to fuel a barbeque) are also expressed in volumes. Coal is an impure solid, so it is measured in units of mass (e.g. kilograms).

To quantify (determine the amount of) the fuel reactants or greenhouse gas products from combustion reactions, you need to be able to calculate the amount (in moles) from mass or volume. These calculations use concepts that you have been taught in previous chapters.

- Enthalpy: $H = \frac{\text{energy }(q)}{\text{mol }(n)}$ or $\frac{\text{energy }(q)}{\text{mass }(m)}$
- Molar volume at SLC ($24.8\,\text{L}\,\text{mol}^{-1}$): $n = \frac{V}{V_m}$
- Density (in g/mL): $d = \frac{\text{mass}}{\text{volume}}$

3.1A Worked example
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3.1A Worked example
Video demonstration

Revision of stoichiometry

Stoichiometric problems can be solved in four steps:

1. Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substances. The known quantity is the one where you can calculate the amount, in mol. The unknown quantity is the substance of interest in the question.
2. Calculate the amount (in mol) of the known quantity of substance present. This may be one calculation or may involve multiple calculations.
3. From the equation, find the molar ratio that states the proportion of known to unknown quantities in the reaction and use it to calculate the amount (in mol) of the unknown substance. This is called a **ratio statement**. It is calculated by dividing the coefficient of the unknown amount by the coefficient of the known substance, then multiplying this value with the amount (in mol) of the unknown substance.
4. Calculate the amount (in mol) of the unknown substance.

Study tip

If the formula is not in the data book, you can develop your own from its units. For example, if density is measured in g/mL, the formula would be:
density = $\frac{\text{mass}}{\text{volume}}$
This can then be rearranged to find mass or volume.

Net mass of major greenhouse gases

Determining the amount of greenhouse gas produced by combustion reactions can help us understand the environmental impact of burning fossil fuels. It can also help quantify how much these gases are contributing to national and international greenhouse gas emissions.

Remember that water is a liquid at SLC. Therefore, water is not included in the total volume of greenhouse gases produced in a combustion reaction. However, it is included in the **net mass** of greenhouse gases produced.

net mass
the specific mass of the desired substance(s), excluding its container

Worked examples 3.1B and 3.1C show you how to calculate the masses and volumes of greenhouse gases produced by combustion reactions.

3.1B WORKED EXAMPLE

CALCULATING THE MASS OF GREENHOUSE GASES

A fuel tank holds 46.00 L of pure liquid octane, which has a density of 0.703 g/mL. Calculate the net mass of greenhouse gases released from the reaction with oxygen at SLC.

Solution

Think	Do
Step 1: Start by writing a balanced chemical equation. This must include states and, for thermochemistry, the enthalpy of the reaction.	$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$ $\Delta H = -10920\ \text{kJ mol}^{-1}$
Step 2: Calculate the mass of octane in the tank using the volume of the full tank and the density of the fuel. Note: The amount, in mol, of octane cannot be calculated using $n = \frac{V}{V_m}$ because this formula applies to gases only, and octane is a liquid.	$V(C_8H_{18}) = 46 \times 1000$ $= 46\,000\ \text{mL}$ $m(C_8H_{18}) = d \times V$ $= 0.703 \times 46\,000$ $= 32\,338\ \text{g}$
Step 3: Calculate the amount, in mol, of the fuel.	$n(C_8H_{18}) = \frac{m}{M}$ $= \frac{32\,338}{114.0}$ $= 283.67\ \text{mol}$
Step 4: Use a ratio statement to calculate the amount, in mol, of CO_2.	$n(CO_2) = n(C_8H_{18}) \times \frac{16}{2}$ $= 283.67 \times 8$ $= 2269.33\ \text{mol}$
Step 5: Calculate the mass of CO_2 formed.	$m(CO_2) = n \times M$ $= 2269.33 \times 44.0$ $= 99\,850.67\ \text{g}$ $= 99.9\ \text{kg}$ (3 sig fig) H_2O is not included because it is a liquid at SLC.

3.1C WORKED EXAMPLE

CALCULATING THE VOLUME OF GREENHOUSE GASES

Calculate the volume of greenhouse gases released into the atmosphere when 50.25 g of methanol is combusted at SLC, according to the following equation:

$$2CH_3OH(g) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$$

Solution

Think	Do
Step 1: Calculate the amount, in mol, of methanol.	$n(CH_3OH) = \frac{m}{M}$ $= \frac{50.25}{32.0}$ $= 1.57\ \text{mol}$
Step 2: Calculate the amount, in mol, of CO_2 using a ratio statement. Note: Only carbon dioxide is a gas at SLC, so only the amount, in mol, of CO_2 needs to be calculated.	$n(CO_2) = n(CH_3OH) \times \frac{2}{2}$ $= 1.57 \times 1$ $= 1.57\ \text{mol}$
Step 3: Calculate the volume of CO_2 using the equation $n = \frac{V}{V_m}$.	$V(CO_2) = n \times V_m$ $= 1.57 \times 24.8$ $= 38.9\ \text{L}$ (3 sig fig) H_2O is not included because it is a liquid at SLC.

Heat energy released

Combustion of fossil fuels also generates heat energy (q). The heat energy released by a volume or mass of fuel can be calculated using the **heat of combustion** of the fuel ($\Delta H°_C$). This is the amount of energy released in the form of heat when one mole or gram of a substance is combusted.

The value of the heat of combustion is the same as that of enthalpy (ΔH), but is reported as positive. The heat of combustion of common fuels will be provided to you in your data book.

For combustion reactions, the following equations are used:

$$q = \Delta H°_C \times n \qquad \text{or} \qquad q = \Delta H°_C \times m$$

where q is energy in kilojoules, $\Delta H°_C$ is heat of combustion in kilojoules per gram or kilojoules per mole, n is the amount in moles and m is the mass in grams.

Have a look at Worked examples 3.1D and 3.1E to see how to calculate energy released by fuels.

heat of combustion
the amount of energy released in the form of heat when one mole or gram of substance is combusted

3.1D Worked example
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3.1D Worked example
Video demonstration

3.1E WORKED EXAMPLE

CALCULATING THE ENERGY IN A MASS OF FUEL

Calculate the energy, in kJ, that must be generated from the combustion of methane in order to produce 1.96 g of water, at SLC.

Solution

Think	Do
Step 1: Start by writing a balanced thermochemical equation. This must include states and enthalpy.	$CH_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$ $\Delta H = -890\ \text{kJ mol}^{-1}$
Step 2: Calculate the amount, in mol, of the known quantity (mass of water).	$n(H_2O) = \frac{m}{M}$ $= \frac{1.96}{18.0}$ $= 0.11\ \text{mol}$
Step 3: Use a ratio statement to calculate the amount, in mol, of the unknown quantity (methane).	$n(CH_4) = n(H_2O)$ $= \frac{0.11}{2}$ $= 0.054\ \text{mol}$
Step 4: Two methods are possible at this point, depending on whether you know the molar heat of combustion (kJ mol^{-1}) or the heat of combustion (kJ g^{-1}) of the substance that is the unknown quantity (methanol) – see your data book. The molar heat of combustion is equal to enthalpy, but without the negative sign.	Calculate the energy (q) using the molar heat of combustion. $q = \Delta H°_C \times n$ $= 890 \times 0.054$ $= 48.5\ \text{kJ}$ (3 sig fig)
	Calculate the energy (q) using the heat of combustion. $m(CH_4) = n \times M$ $= 0.054 \times 16$ $= 0.87\ \text{g}$ $q = \Delta H°_C \times m$ $= 55.6 \times 0.87$ $= 48.4\ \text{kJ}$ (3 sig fig)

3.1 Skill drill
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If you feel confident, try Challenge 3.1, or practise your data analysis and evaluation skills in Skill drill 3.1.

3.1 CHALLENGE

Volume of gases at non-standard conditions

Calculate the volume of greenhouse gases released into the atmosphere when 5.00 kg of ethyne (C_2H_2) is combusted at 200°C and 105 kPa.

1 Calculate the amount, in mol, of ethyne.
2 Calculate the amount, in mol, of the gases produced in the reaction.
3 Calculate the volume and the net volume of gases produced.

3.1 CHECK YOUR LEARNING

Describe and explain

1 Identify the state of water at SLC.
2 Identify which greenhouse gas is produced in the highest quantity by combustion reactions of hydrocarbons at SLC. Explain the environmental implications of this.

Apply, analyse and compare

3 Balance the following combustion reactions:
  a $C_3H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
  b $Al(s) + O_2(g) \rightarrow Al_2O_3(s)$
  c $N_2(g) + O_2(g) \rightarrow NO(g)$
  d $C_4H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
  e $C_6H_{12}O_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
4 Calculate the energy required to produce 1.00 kg of water from the complete combustion of octane.
5 Calculate the mass of propane required to produce 500 g of water at SLC.
6 Calculate the energy released when a 48.0 L liquid tank of gaseous propane undergoes complete combustion at SLC (density = 0.493 g mL^{-1}).
7 A butane torch is typically a 300 mL cylinder (density = 0.573 g mL^{-1}). Calculate the amount of energy that a butane torch contains at SLC.
8 An acetylene (C_2H_2) torch is fuelled by a 7.20 L liquid cylinder (density = 1.18 g L^{-1}). Acetylene burns at over 1800°C in pure oxygen, which makes it ideal for use in a welding torch, because it will melt most metals. In air, however, it burns at 1400°C. If air contains 20% oxygen, calculate the volume of air required to combust the acetylene in the tank at SLC.

3.2

Specific heat capacity

KEY IDEAS

In this topic, you will learn that:

- the specific heat capacity of water is the energy that it takes to raise the temperature of 1 g of water by 1°C
- the change in temperature can be used to measure energy released from combustion of fuel and food.

The energy released by the combustion of fuels is used at home and in many industries. If the energy is used to heat a specific volume of water, such as when you use a gas stove to boil water for cooking, the change in temperature of the water can be measured. We use this to calculate the energy required to heat the water (the energy generated by the fuel when it combusts).

Using specific heat capacity to calculate energy

specific heat capacity
a measure of the amount of heat energy in joules (J) or kilojoules (kJ) that it takes to increase the temperature of 1 g of a substance by 1°C

Specific heat capacity (SHC) is a measure of the amount of heat energy, in joules (J), that it takes to increase the temperature of 1 g of a substance by 1°C. The specific heat capacity of water, under ideal conditions, is $4.18\,\mathrm{J\,g^{-1}\,K^{-1}}$ or $4.18\,\mathrm{kJ\,kg^{-1}\,°C^{-1}}$. Remember from Unit 2 that water has a relatively high SHC due to its ability to form strong hydrogen bonds with itself. SHC can be used to calculate the heat energy released from a combustion reaction using the following equation:

$$q = m \times c \times \Delta T$$

where q is heat energy in joules, m is mass of the water in grams, c is the specific heat capacity of water and ΔT is the change in temperature in kelvin or degrees Celsius.

Units of specific heat capacity

SHC is measured in joules per gram per kelvin ($\mathrm{J\,g^{-1}\,K^{-1}}$) or kilojoules per kilogram per kelvin ($\mathrm{kJ\,kg^{-1}\,K^{-1}}$). Values can also be expressed per degrees Celsius (°C) instead of kelvin. This provides flexibility in the values that you can use to calculate energy.

- If the mass of water is in kilograms (kg), then the unit of energy is kilojoules (kJ).
- If the mass of water is in grams (g), then the unit of energy is joules (J).

As heat of combustion is measured in $\mathrm{kJ\,mol^{-1}}$, it is useful to convert the mass of water to kg and calculate energy directly in kJ (see Worked example 3.2D) rather than calculating it in J and then converting it to kJ.

Study tip

Remember that $q = m \times c \times \Delta T$ refers specifically to water if $c = 4.18\,\mathrm{J\,g^{-1}\,K^{-1}}$ or $4.18\,\mathrm{kJ\,kg^{-1}\,K^{-1}}$. q represents the energy required to heat a mass of water by a specific temperature.

3.2A WORKED EXAMPLE

CALCULATING THE ENERGY REQUIRED TO HEAT WATER

200 g of water was heated in a beaker. The initial temperature of the water was 16°C and the final temperature was 62°C. Calculate the amount of energy (kJ) required to heat the water.

Solution

Think	Do
Step 1: Calculate the energy in J using $q = m \times c \times \Delta T$.	$q = m \times c \times \Delta T$ $= 200 \times 4.18 \times (62 - 16)$ $= 38\,456\,\mathrm{J}$
Step 2: Convert J to kJ.	$38\,456 \div 1000 = 38\,\mathrm{kJ}$ (2 sig fig)

3.2B WORKED EXAMPLE

CALCULATING THE ENERGY RELEASED FROM THE COMBUSTION OF FUEL

A butane burner was used to heat 600 g of water in a beaker. The initial temperature of the water was 21°C and the final temperature was 42°C. Assuming that 20% of the heat energy was lost to the environment, calculate the total energy (kJ) released from the burner to heat the water.

Solution

Think	Do
Step 1: Calculate the energy in kJ.	$q = m \times c \times \Delta T$ $= 600 \times 4.18 \times (42 - 21)$ $= 52\,668$ J $52\,668 \div 1000 = 52.67$ kJ
Step 2: Assuming that 52.67 kJ is 80%, use the percentage to calculate the energy in 100%.	If 80% of the energy is absorbed by the water, 52.67 kJ is 80% of the energy. Therefore, $\frac{52.67}{80} = 0.658$ kJ is the equivalent of 1%. Therefore, $0.658 \times 100 = 66$ kJ (2 sig fig)

3.2C WORKED EXAMPLE

CALCULATING THE ENERGY RELEASED FROM COMBUSTION OF FUEL

A can of propane is used to raise the temperature of 900 g of water by 16°C. Calculate the energy of the fuel in kilojoules (kJ).

Solution

If a question asks you to calculate energy in kJ, it may be easier to convert the mass to kg and calculate energy in kJ directly (method 1). If you choose to keep the mass of water in g, you must remember to convert it to kJ later (method 2).

Think	Do
Method 1	
Step 1: Convert the mass of water to kg.	$900 \div 1000 = 0.900$ kg
Step 2: Use the SHC of water in $\text{kJ kg}^{-1}\text{K}^{-1}$ to calculate energy in kJ.	$q = m \times c \times \Delta T$ $= 0.900 \times 4.18 \times 16$ $= 60$ kJ (2 sig fig)
Method 2	
Step 1: Use the SHC of water in $\text{J g}^{-1}\text{K}^{-1}$ to calculate energy in J.	$q = m \times c \times \Delta T$ $= 900 \times 4.18 \times 16$ $= 60\,192$ J
Step 2: Convert the energy from J to kJ.	$60\,192 \div 1000 = 60$ kJ (2 sig fig)

Distinguishing between heat and temperature

It is important to understand the distinction between heat and temperature. Heat is a form of energy referred to as *thermal energy*. Temperature is a measure of the average kinetic energy of the particles. Therefore, temperature is a measure of thermal (heat) energy. Explore the energy released from nitroglycerin and trinitrotoluene (TNT) in Real-world chemistry 3.2.

3.2 Real-world chemistry
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Measuring energy released from food products

As long as the change in the temperature of water can be measured, this equation can also be used to calculate the energy generated from the combustion of other materials, such as food.

As foods are not pure substances, but a mixture of the ingredients and biomolecules that are within them, their heat of combustion cannot be calculated in kJ mol^{-1}. Instead, any energy that is referenced on food packaging must be calculated in kJ per gram of the food.

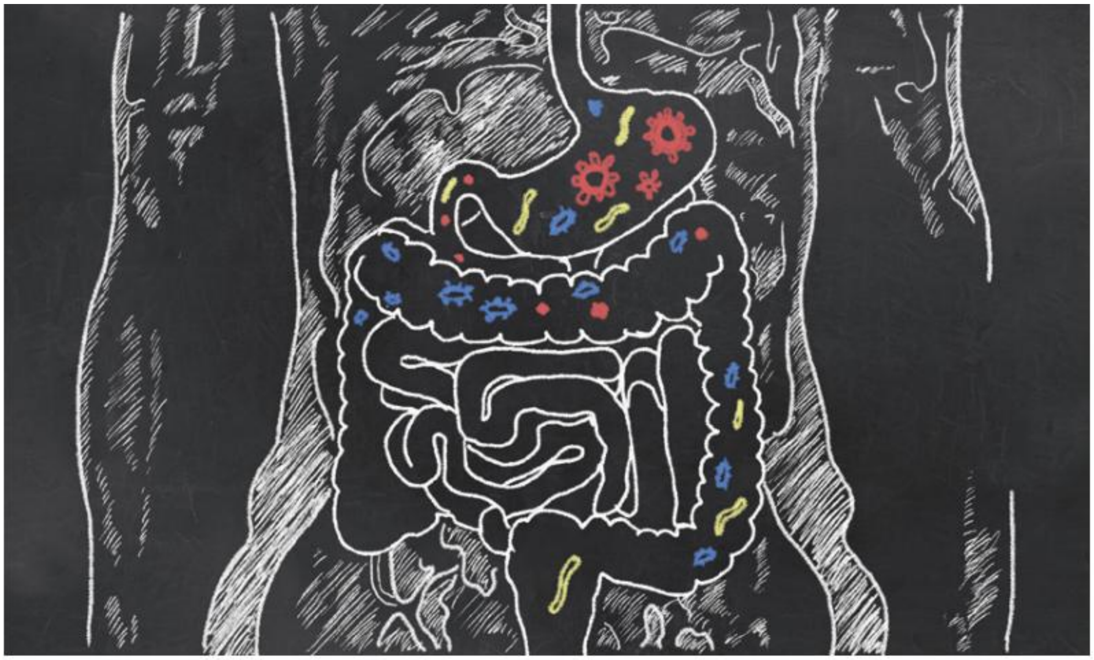


FIGURE 1 In the gastrointestinal tract, food is combusted to release energy.

3.2 CHECK YOUR LEARNING

Describe and explain

1 Explain what is measured by specific heat capacity.

2 Explain what must be done in a calculation to use specific heat capacity in units of $\text{kJ kg}^{-1}\text{K}^{-1}$ instead of $\text{J g}^{-1}\text{K}^{-1}$.

3 Identify whether a denser substance would have a higher or lower specific heat capacity. Explain your reasoning.

Apply, analyse and compare

4 Determine the amount of energy required to heat 72 g of water, if the initial temperature of the water was 296 K and the final temperature was 54°C.

5 Calculate the final temperature of 250 g of water, if its initial temperature is 18.2°C and 20 kJ of energy is used to heat it.

6 900 J of energy was applied to a specific mass of water to raise its temperature by 40°C. Calculate the mass of water that the energy was applied to.

Design and discuss

7 Discuss why solid ice would have a different specific heat capacity than liquid water. State whether it would be higher or lower. Justify your response.

8 Discuss the property of water that gives it a relatively high specific heat capacity and explain how it does this. Hint: Why are water molecules held relatively closely together?

3.3 The principles of solution calorimetry

KEY IDEAS

In this topic, you will learn that:

- calorimeters are used to measure the temperature change of solutions with minimal heat loss to the environment
- calorimeters must be calibrated to determine the amount of energy that it takes to increase the temperature of a solution by 1°C
- calorimeters can be used to determine the energy released by different fuels and food products.

In Topics 3.1 and 3.2, you learned three main points:

1 Heat energy is released in a combustion reaction.
2 The energy released can be used to heat water.
3 You can calculate the energy released depending on the change in water temperature.

In this topic, you will learn about how we can experimentally determine the amount of energy released from reactions in the laboratory, using a technique called solution calorimetry. The development of this technology was especially important for chemical engineers and industrial chemists in the latter half of the 20th century. This helped to make processes more efficient and safer, and reduce waste.

Principles of calorimetry

When you boil water on a stove, you are burning (combusting) fuel to heat the pot of water. The change in temperature of the known volume of water can be measured to determine how much energy is being supplied by the fuel. In the same way, you can burn food (a form of fuel) to heat a known volume of water in a can and measure the change in water temperature, to then determine the energy released from the food.

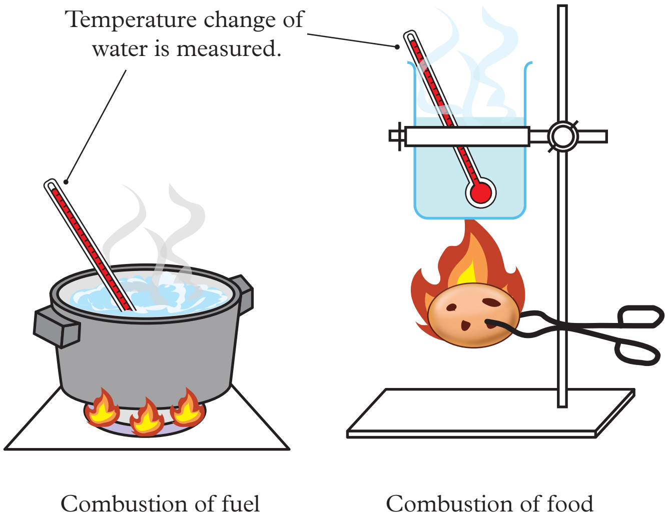


FIGURE 1 Energy released from combustion of fuel and food can be determined by measuring the change in the temperature of water.

3.3
PRACTICAL: CONTROLLED EXPERIMENT

How can we measure the enthalpy of a reaction? Go to page **503**.

This is the basic principle of **solution calorimetry** – heat energy generated from a reaction is used to heat a known mass of water; you measure the change in temperature (ΔT) of the water; you use this ΔT value to calculate the energy released from the reaction.

Heat energy is also released from thermochemical solutions. Therefore, you can also use solution calorimetry to measure the change in temperature of water – and then energy released – after dissolving a known quantity of substance in water (**dissolution**) or reacting a known quantity of substance with another chemical in an **aqueous reaction**.

However, there is one major problem with calorimetry conducted using the apparatuses in Figure 1: you will lose a lot of heat energy from the water to the environment. This can affect the accuracy and validity of your results. To make sure your experiment is valid, you will need to conduct a calibration; even better, you can improve accuracy by using a specific device called a **calorimeter**.

solution calorimetry
an analytical technique in which the energy generated by a sample is determined by measuring the change in the temperature of water

dissolution
the process of dissolving a solute into a solvent (typically water) to create a solution

aqueous reaction
the reaction of two solutions that form aqueous products

calorimeter
a device that measures the change in temperature of water when an endothermic or exothermic reaction takes place

A calorimeter

A solution calorimeter is any device designed to measure to a high level of accuracy the rise or fall in temperature of a quantity of solution during an endothermic or exothermic reaction.

The calorimeter in Figure 2 contains an insulated container, which minimises heat loss to the surrounding environment. The solution is placed inside the container with a thermometer, which measures the change in temperature. The stirrer ensures even distribution of thermal energy in the container. An even more basic type of calorimeter is shown in Figure 3.

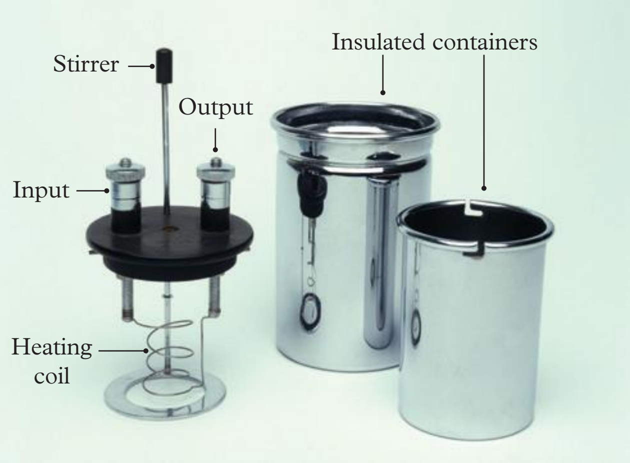


FIGURE 2 The parts of a solution calorimeter

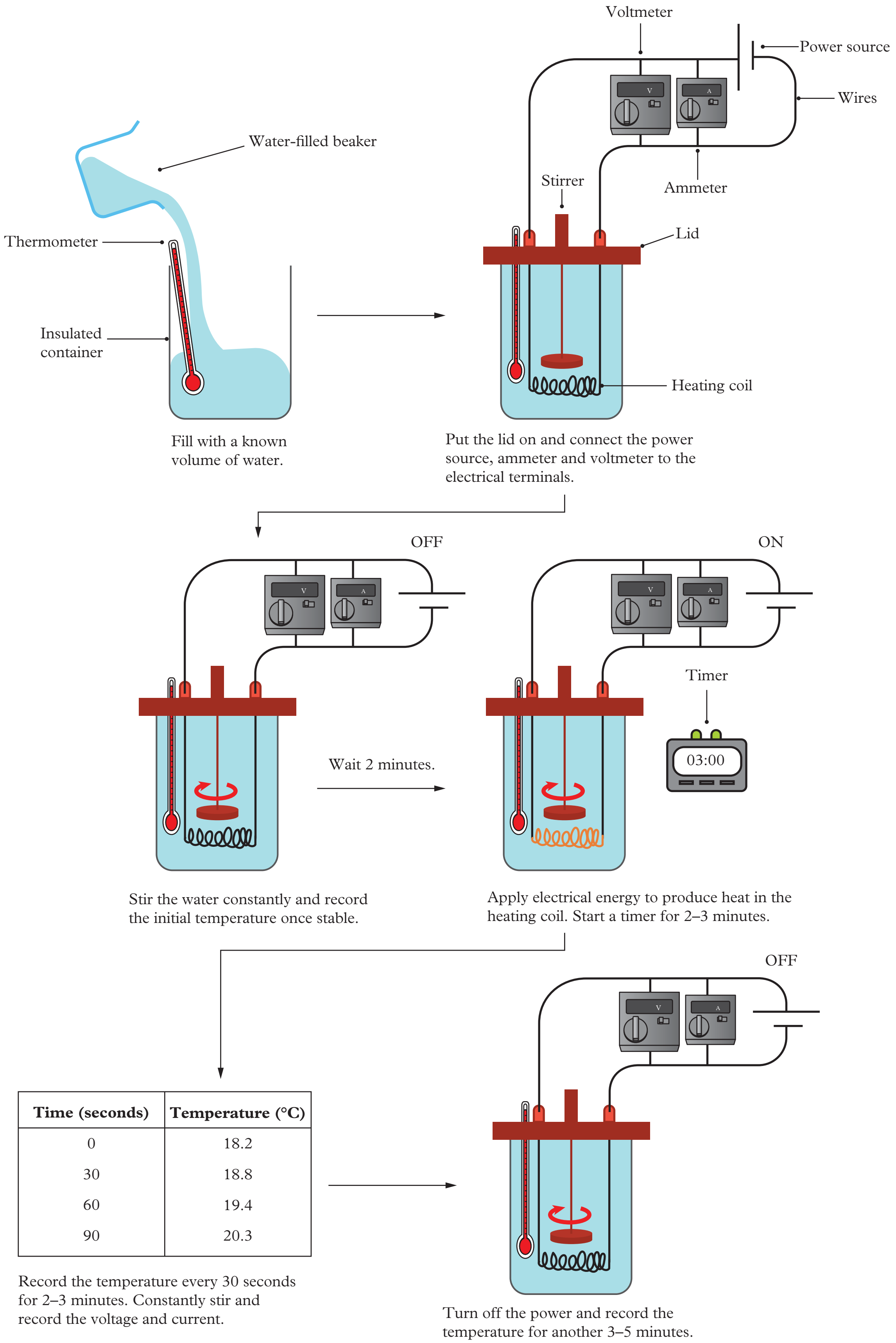
FIGURE 3 A very basic calorimeter that you might use in the laboratory

Calibrating a calorimeter

A calorimeter must be calibrated before use. Calibration (also called electrical calibration) determines the amount of energy that it takes to change the temperature of the solution in the calorimeter by 1°C. This is called the **calibration factor** (CF), and it is used in energy calculations.

All calorimeters have slightly different calibration factors, making it essential to calibrate every calorimeter before it is used. The steps for calibration are shown in Figure 4, on the next page.

calibration factor
the amount of energy that it takes to increase the temperature of an aqueous substance in a calorimeter by 1°C. This value is used to calibrate the calorimeter.



Fill with a known volume of water.

Put the lid on and connect the power source, ammeter and voltmeter to the electrical terminals.

Stir the water constantly and record the initial temperature once stable.

Apply electrical energy to produce heat in the heating coil. Start a timer for 2–3 minutes.

Time (seconds)	Temperature (°C)
0	18.2
30	18.8
60	19.4
90	20.3

Record the temperature every 30 seconds for 2–3 minutes. Constantly stir and record the voltage and current.

Turn off the power and record the temperature for another 3–5 minutes.

FIGURE 4 Steps for calibrating a calorimeter

Determining the efficiency of a calorimeter

Once you have your results, you can plot time on the *x*-axis and temperature on the *y*-axis. This makes a **calibration curve** (Figure 5). The calibration curve is important to make sure the calorimeter is working efficiently. This will affect the validity of your results.

You make this assessment by looking at the shape of the curve – particularly the region after the power is turned off.

calibration curve (in calorimetry) a graph of temperature versus time used to determine the efficiency of a calorimeter

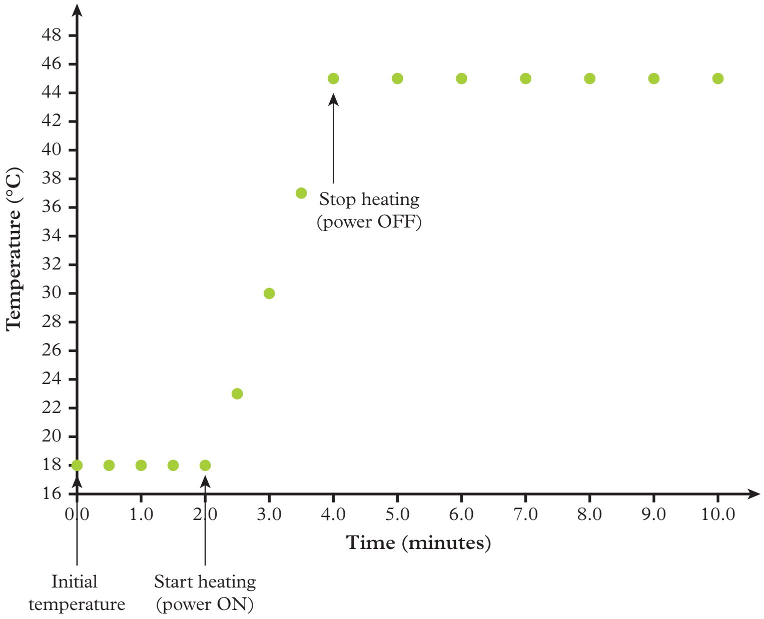


FIGURE 5 A calibration curve for a solution calorimeter

Efficient calorimeters

Efficient calorimeters can retain the heat applied to them and will not lose it to the environment (Figure 6). Calculating the change in temperature is still necessary, but the back extrapolation is far simpler due to the lack of heat loss.

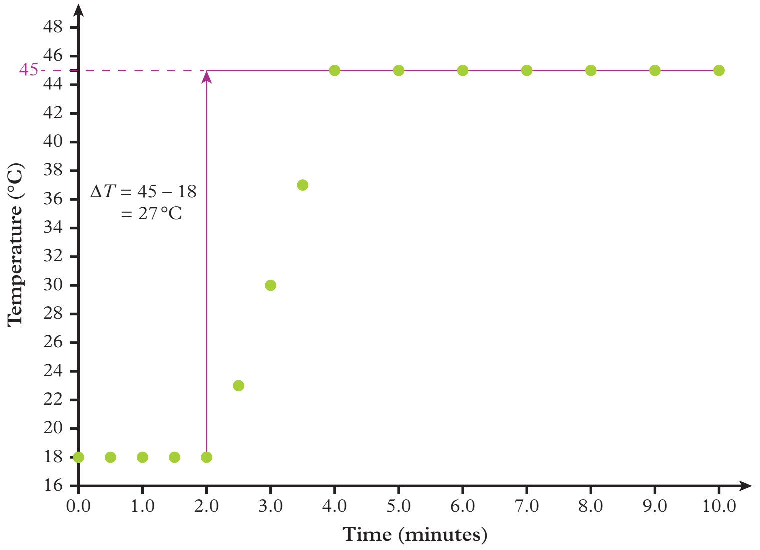


FIGURE 6 Determining the change in temperature (ΔT) using an efficient calorimeter

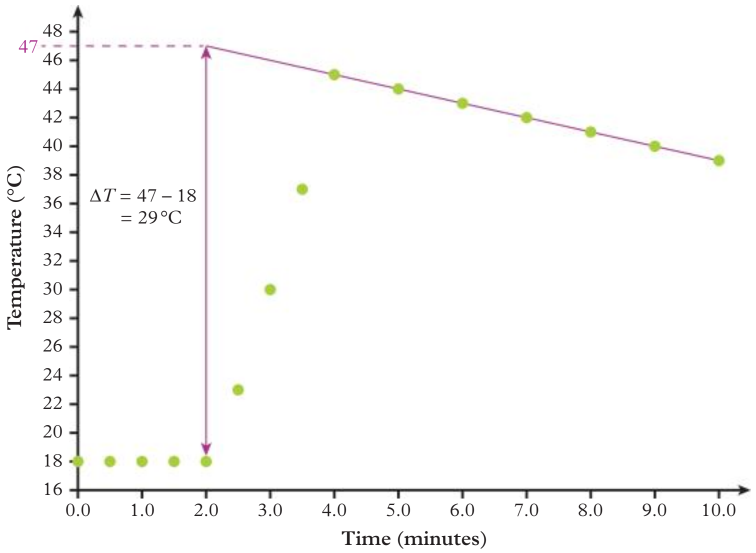


FIGURE 7 Determining the change in temperature (ΔT) using an inefficient calorimeter

Inefficient calorimeters

A calorimeter is considered inefficient if it loses heat to the environment. The initial temperature of the water may be constant (as it stabilises to room temperature), but the water will lose heat energy to the environment as it is warmed by the heating coil.

A key sign that a calorimeter is losing heat to the environment is that the temperature of the water in the calorimeter decreases once the power source is turned off, which can be seen by the downward slope in the calibration curve (Figure 7).

If the calorimeter loses heat energy, then the final temperature of the calorimeter is inaccurate. For this reason, the data must be extrapolated backward to determine the final temperature, assuming no heat loss.

The change in temperature of the calorimeter is calculated as:

$$\Delta T = \text{final temperature} - \text{initial temperature}$$

Because ΔT is the difference in temperature, it does not matter whether it is measured in degrees Celsius (°C) or in kelvin (K).

An example of an inefficient calorimeter would be Figure 1, where a pot of water is heated using combusted fuel (from a stove), or a beaker or can of water is heated using energy from a burning cookie. You can still extrapolate the data from this kind of simple calorimeter and use it to determine energy released by the fuel or food. However, you must be careful when analysing data from an inefficient calorimeter because significant loss of heat from the calorimeter may make your experiment invalid.

Study tip

An inefficient calorimeter is still usable, as long as you conduct a calibration beforehand, but be careful if the heat loss is too significant.

Calculating the calibration factor (CF)

To calculate the CF of a calorimeter, the calibration curve (Figure 4) is used to determine the change in temperature.

The energy (E) applied to the calorimeter is calculated as:

$$E = V \times I \times t$$

where E is energy in joules, V is voltage in volts, I is current in amps, and t is time in seconds.

CF is calculated as energy per degree of temperature. Therefore, it is measured in J °C^{-1} or J K^{-1}, and calculated using:

$$\text{CF} = \frac{E}{\Delta T}$$

Alternatively, the two equations can be combined to give the following equation:

$$\text{CF} = \frac{V \times I \times t}{\Delta T}$$

See how to calculate the calibration factor for a calorimeter in Worked example 3.3A.

Once your calorimeter has been calibrated, it is ready to be used. The water used to calibrate the calorimeter can be disposed of and the power source with the ammeter and voltmeter are no longer necessary, so they can be disconnected.

3.3A Worked example
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3.3A Worked example
Video demonstration

Measuring the energy released from solutions

FIGURE 8 The energy or heat of dissolution can be measured.

It can be assumed that aqueous solutions have the same specific heat capacity as water, since aqueous solutions are solutes that have been dissolved in a water solvent. Therefore, the energy of aqueous solutions can be calculated in the same way.

If the calibration factor of the calorimeter is known, then the energy of the reaction can be calculated as:

$$E = \text{CF} \times \Delta T$$

where E is energy in joules, CF is the calibration factor, and ΔT is the change in temperature during the reaction.

The energy of a solution is better expressed as **molar heat of the solution** (kJ mol^{-1}).

When one mole of a solid substance dissolves into water to form an aqueous solution (e.g. when solid NaCl dissolves in water), this is called the **molar heat of dissolution**. See how this is calculated in Worked example 3.3B.

molar heat of solution
the amount of energy absorbed or released by a solution

molar heat of dissolution
the enthalpy of a solution when a solid substance dissolves in water

When two aqueous solutions react, such as in a neutralisation reaction between an acid and a base, this is called the **molar heat of neutralisation**. One of the aqueous reactants is placed in the calorimeter and allowed to reach a constant temperature, which is recorded as the initial temperature. The second reactant (solid or aqueous) is added with constant stirring and allowed to react. The maximum/minimum temperature is recorded and the change in temperature is calculated. Have a look at how to calculate molar heat of neutralisation in Worked example 3.3C.

3.3B Worked example
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3.3B Worked example
Video demonstration

This is different from the enthalpy (ΔH) of a solution (in kJ mol^{-1}), which takes into account the nature of the reaction, i.e. if it gains or loses energy.

Remember:

- exothermic reactions will cause an increase in the temperature of the surroundings (water), due to a loss of energy from the reactants. Therefore, the enthalpy of the solution will have a negative value
- endothermic reactions will cause a decrease in the temperature of the surroundings (water), due to a gain of energy by the products. Therefore, the enthalpy of the solution will have a positive value.

molar heat of neutralisation
the enthalpy of a solution when an acid reacts with a base

Measuring the energy released from food and fuel

As you saw in Figure 1 at the beginning of this topic, the energy or heat of combustion for food and fuel can also be determined by calorimetry. As long as calibration is conducted beforehand, the experiment is considered valid.

You would ideally use an insulated, enclosed container, instead of the beaker shown in the image. However, the combustion of food and fuel requires an ignition source and an excess supply of oxygen, which a solution calorimeter does not accommodate. Therefore, an alternative device called the bomb calorimeter is required.

3.3C Worked example
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3.3C Worked example
Video demonstration

The principles of bomb calorimetry are the same as those of solution calorimetry. Bomb calorimetry is not included in the VCE Chemistry study design, but you can apply your understanding of the principles of calorimetry by attempting Challenge 3.3.

A special device called a spirit burner can also be used to measure the heat of combustion of a fuel. You will explore this in the next topic.

3.3 CHALLENGE

Bomb calorimeter

In a bomb calorimeter, a sample of solid is contained within a reaction capsule. An electric current is used to ignite the sample, which is fed with a constant supply of excess oxygen to cause the whole sample to react. The temperature change of the water is recorded, and the energy of the food is calculated.

A bomb calorimeter was used to combust a sample of Cheezels to compare the energy of the Cheezels to the nutritional value on the back of the packaging. 0.5 g of Cheezels were combusted in excess oxygen to compare to the experimental heat of combustion of 2270 kJ per 100 g. Assuming that the bomb calorimeter holds a mass of 300 g of water, calculate the temperature rise of the water in the bomb calorimeter.

Practise your skills using Skill drill 3.3 to make sure you understand the potential errors involved in calorimetry.

3.3 SKILL DRILL

Errors in calorimetry

Key science skill: Analyse and evaluate data and investigation methods

If a calorimeter is inefficient, it loses heat energy to the environment. This is an error that cannot be resolved by repeating the experiment, because the issue is with the equipment itself.

Practise your skills

1 Identify what type of error this is. Explain your reasoning.

2 Explain how this type of error can be minimised.

3 A calorimeter had a calibration factor of $12\,J\,°C^{-1}$ when it was first purchased. Frequent use has resulted in the degradation of the insulated container. In an experiment to determine the energy released in the combustion of a solution, the calibration factor of $12\,J\,°C^{-1}$ was used.

  a Explain whether the energy released by the solution would be higher or lower than its theoretical value.

  b Evaluate the accuracy and validity of the data.

Need help analysing and evaluating data? See Topic 1.8 (page 24).

3.3 CHECK YOUR LEARNING

Describe and explain

1 Explain how a calorimeter is able to measure the enthalpy of a solution.

2 Explain why a calorimeter is insulated. Describe what would happen if it was not insulated.

3 Explain why the combustion of solids, liquids and gases cannot be conducted in a solution calorimeter.

Apply, analyse and compare

4 To calibrate a calorimeter, a student let the 150 g of water in the calorimeter reach a constant temperature of 18.1°C for 3 minutes before turning on a power pack. The power pack applied 6.0 V of electrical energy for 2.5 minutes, generating a current of 2.1 A. The following data was obtained.

Time (mins)	Temperature (°C)
0.0	18.1
0.5	24.5
1.0	31.6
1.5	39.2
2.0	46.3
2.5	53.2
3.0	52.7
3.5	52.2
4.0	51.6
4.5	51.1
5.0	50.2

a Construct a calibration graph for the data.

b Determine the change in temperature from the graph.

c Calculate the calibration factor of the calorimeter.

5 To calibrate a calorimeter, a student let 100 g of water in the calorimeter reach a constant temperature of 16.5°C for 2.5 minutes before turning on a power pack. The power pack applied 4.0 V of electrical energy for 2.0 minutes, generating a current of 1.5 A. The following data was obtained.

Time (mins)	Temperature (°C)
0.0	16.5
0.5	20.3
1.0	24.1
1.5	28.9
2.0	32.4
2.5	32.0
3.0	31.5
3.5	31.1
4.0	30.6
4.5	29.9
5.0	29.2

a Construct a calibration graph for the data.

b Determine the change in temperature from the graph.

c Calculate the calibration factor of the calorimeter.

6 A calorimeter is calibrated at 20 J °C^{-1}. Calculate the enthalpy (kJ mol^{-1}) of dissolution of 0.50 g of KNO_3, if the initial temperature is 19.5°C and the final temperature is 11.0°C. Identify whether this reaction is exothermic or endothermic.

7 A 1.20 g sample of potassium iodide (KI) is dissolved in a calorimeter filled with 250 g of water. The calorimeter has been previously calibrated, and its calibration factor was found to be 32 J °C^{-1}. The starting temperature of the water was 21.3°C and the final temperature was 16.8°C. Calculate the enthalpy of KI in kJ mol^{-1}.

8 A student intends to calculate the enthalpy of neutralisation of 50 mL of 1.5 M hydrochloric acid (HCl). The student reacts the HCl with 50 mL of 1.5 M sodium hydroxide (NaOH) in a calorimeter that has a calibration factor of 16 J °C^{-1}. The initial temperature of the HCl solution was 15.2°C and its final temperature was 77.1°C. Calculate the enthalpy of neutralisation of HCl in kJ mol^{-1}.

Design and discuss

9 Discuss why a calorimeter must be calibrated before it is used.

10 A student obtains the results for the calibration of the calorimeter in Figure 6. Instead of extrapolating the final temperature, the student concludes that it is 45°C (rather than 47°C). After this, they conduct several reactions and measure the enthalpy of these reactions. Discuss the effect of this error on the calculated energy released from the reaction. Explain whether this is a random or systematic error.

3.4 Energy from fuels and food

KEY IDEAS

In this topic, you will learn that:

- the energy released by combustion of blended fuels must be calculated in kilojoules per gram
- the energy of food can be calculated based on the amount of proteins, fats and carbohydrates in the food
- energy transformation efficiency of combustion reactions can be calculated by comparing the experimental energy to the theoretical energy.

In Topic 3.2, you learned how to calculate the heat of combustion of food and fuel, using specific heat capacity. Then, in Topic 3.3, you learned how to determine the energy released from a solution (heat or enthalpy of solution) in $kJ\,mol^{-1}$.

Energy from food and fuel can also be calculated, but, in many cases, cannot be expressed per mol. This is because food and fuel are often impure mixtures. Therefore, they are given as $kJ\,g^{-1}$.

Recall the equation:

$$q = \Delta H^\circ_C \times m$$

where q is the energy in kJ, ΔH°_C is the heat of combustion in $kJ\,g^{-1}$ and m is the mass.

Study tip

The fuels looked at in Topics 3.1 and 3.2 have all been pure fuels. They have known molar masses, so a mole amount can be calculated. Other fuels, such as diesel, kerosene and natural gas, are mixtures of different molecules. For this reason, their heat of combustion cannot be calculated in $kJ\,mol^{-1}$ and must instead be calculated in $kJ\,g^{-1}$.

Energy from fuels

There are a few impure or blended fuels that we use regularly to generate the energy we rely on. A list of these fuels, including their states and heats of combustion (in $kJ\,g^{-1}$), is found in Table 1.

TABLE 1 Heat of combustion of common blended fuels (at 25°C and 100 kPa)

Blended fuel	State	Heat of combustion ($kJ\,g^{-1}$)
Kerosene	liquid	46.2
Diesel	liquid	45.0
Natural gas	gas	54.0
Black coal (anthracite)	solid	32.5
Brown coal (lignite)	solid	15.0
Wood (moisture free)	solid	21.7
Wood (firewood)	solid	21.2
Peat (dry)	solid	15.0
Peat (damp)	solid	6.00

3.4A Worked example
Find me in your obook pro

3.4A Worked example
Video demonstration

Because these fuels are mixtures, they have no chemical formula and are not placed into balanced chemical equations, so stoichiometric ratios cannot be obtained.

The energy released when an impure substance combusts is calculated using the same equation used earlier for heat of combustion (Worked example 3.4A).

Calculating the heat of combustion

For fuels that are liquids, a **spirit burner** can be used to calculate the heat of combustion. This is calculated in a similar way to the heat of solution.

spirit burner
a vessel with a wick that can hold liquid fuels and combust them in the presence of oxygen in the air

Figure 1 shows a spirit burner, which efficiently burns a liquid fuel while keeping it contained. The mass of the spirit burner is recorded before and after the experiment to determine the mass of the fuel used to heat the water above the burner. A set volume of water is contained within the beaker, or metal can, with a thermometer and stirrer.

This set-up might look similar to what you saw in Topic 3.3. The principles are the same as those of calorimetry – the reaction heats up water, temperature change is recorded, and a calculation is done to determine the energy released.

Because the mass of the fuel and the temperature rise of the water are known, a $\Delta H°_C$ value can be calculated. Worked examples 3.4B and 3.4C show how to calculate $\Delta H°_C$ for pure and blended fuels, respectively.

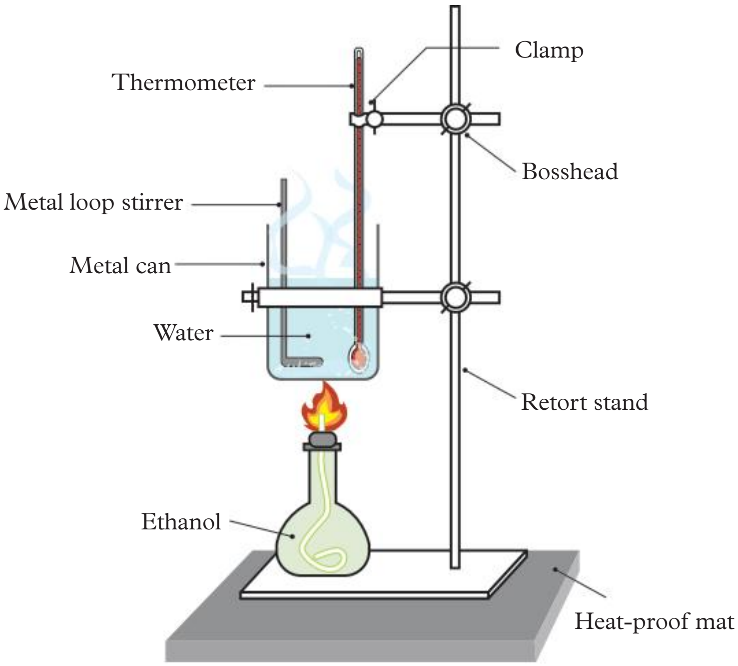

FIGURE 1 A spirit burner records the enthalpy change in a combustion reaction.

3.4B WORKED EXAMPLE

CALCULATING THE HEAT OF COMBUSTION OF A PURE FUEL

A 4.60 g sample of ethanol (C_2H_5OH) was burnt using a spirit burner. The energy produced by the flame was used to heat a beaker filled with 150 g of water. The starting temperature of the water was 18.00°C and the final temperature was 68.00°C. Calculate the heat of combustion for, in kJ mol^{-1}, of ethanol.

Solution

Think	Do
Step 1: Calculate the energy absorbed by the water.	$q = m \times c \times \Delta T$ $= 150 \times 4.18 \times (68 - 18)$ $= 31\,350\text{ J}$ $31\,350 \div 1000 = 31.35\text{ kJ}$
Step 2: This is the energy for 4.60 g of ethanol. Convert mass to moles and calculate the $\Delta H°_C$, in kJ mol^{-1}, of ethanol.	$n(C_3H_8) = \frac{m}{M}$ $= \frac{4.60}{46.0}$ $= 0.10\text{ mol}$ $\Delta H°_C = \frac{q}{n}$ $= \frac{31.35}{0.10}$ $= 307\text{ kJ mol}^{-1}$ (3 sig fig)

3.4C Worked example
Find me in your obook pro

3.4C Worked example
Video demonstration

Study tip

Only alcohol-based fuels (e.g. methanol, ethanol) can be burned in a spirit burner.

Efficiency of combustion reactions

Energy efficiency is important to consider when combusting fuels to calculate their heat of combustion. Using a spirit burner as an example, if energy is lost to the environment instead of being used to heat the water, the enthalpy calculated in the combustion reaction will be lower than expected. This is because the same mass of fuel is being used to heat the water to a lower temperature, and therefore the experimental enthalpy will be lower than the theoretical enthalpy from your data book.

This is referred to as the energy transformation efficiency. It is calculated by dividing the experimental energy by the theoretical energy and expressing the answer as a percentage. This value is the percentage of the energy in the combusted fuel that has been transferred to the water. The remaining percentage is the energy that has been lost, typically to the environment.

$$\%\text{efficiency} = \frac{\text{experimental energy or heat of combustion}}{\text{theoretical energy or heat of combustion}} \times 100$$

Worked example 3.4D shows you how to calculate energy efficiency. Once you are confident with this calculation, have a look at how to use energy efficiency to calculate change in temperature (Worked example 3.4E).

3.4D WORKED EXAMPLE

CALCULATING ENERGY EFFICIENCY

Refer to the data in Worked examples 3.4B and 3.4C to answer the following questions.

a Calculate the energy efficiency of both experiments.

b Determine which experiment was more energy efficient.

c Determine which experiment lost a greater amount of energy to the environment.

Solution

Think	Do
Step 1: Find the theoretical heat of combustion values in your data book.	Ethanol: 1360 kJ mol^{-1} Kerosene: 46.2 kJ g^{-1}
Step 2: Calculate the energy efficiency of the two experiments.	**a** $\%\text{efficiency(ethanol)} = \frac{\text{exp enthalpy}}{\text{theor enthalpy}} \times \frac{100}{1}$ $= \frac{-307}{-1360} \times 100$ $= 23\%$ $\%\text{efficiency(kerosene)} = \frac{\text{exp enthalpy}}{\text{theor enthalpy}} \times \frac{100}{1}$ $= \frac{-29.0}{-46.2} \times 100$ $= 63\%$
Step 3: Compare the efficiency in the two experiments.	**b** The kerosene experiment was more energy efficient. **c** The ethanol experiment lost more energy as heat to the surrounding environment.

3.4E Worked example
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3.4E Worked example
Video demonstration

Energy from food

Foods contain three main biomolecules: fats/oils, proteins and carbohydrates. Each of these biomolecules contains an amount of energy per gram of biomolecule. As you learnt in Chapter 2, when food is eaten, combustion of the food within your digestive system generates chemical energy for bodily processes, such as muscle and organ function. Table 2 shows the heat of combustion for these biomolecules.

TABLE 2 Heat of combustion of biomolecules

Food	Heat of combustion ($kJ\,g^{-1}$)
Fats and oils	37
Protein	17
Carbohydrate	16

The recommended energy intake for an adult is 8700 kJ per day. It is important that you understand food packaging and nutritional information, so that you can make informed decisions about your food intake. The panel on the left-hand side of Figure 2 shows the nutritional information on a packet of biscuits.

nutrition information

NUTRITION INFORMATION (Average)
SERVINGS PER PACKAGE: 10
SERVING SIZE: 25 g (6 biscuits)

	QUANTITY PER SERVING	% DAILY INTAKE* (PER SERVING)	QUANTITY PER 100 g
ENERGY	493 kJ	5.6%	1970 kJ
PROTEIN	2.0 g	3.9%	7.9 g
FAT, TOTAL	5.0 g	7.2%	20.0 g
- SATURATED	1.0 g	4.1%	3.9 g
CARBOHYDRATE	15.6 g	5.0%	62.6 g
- SUGARS	0.3 g	0.4%	1.3 g
SODIUM	212 mg	9.2%	848 g

*BASED ON AN AVERAGE ADULT DIET OF 8700 KJ.
ALL VALUES CONSIDERED AVERAGES UNLESS OTHERWISE INDICATED

nutrition information

NUTRITION INFORMATION (Average)
SERVINGS PER PACK: 10
SERVING SIZE: 200 mL

	QUANTITY PER SERVING	% DAILY INTAKE* (PER SERVING)	QUANTITY PER 100 mL
ENERGY	306 kJ	4%	153 kJ
PROTEIN	1.2 g	2%	0.6 g
FAT, TOTAL	<1 g	<1%	<1 g
- SATURATED	0 g	0%	0 g
CARBOHYDRATE	18.0 g	6%	9.0 g
- SUGARS	16.6 g	18%	8.3 g
DIETARY FIBRE	0.4 g	1%	0.2 g
SODIUM	16 mg	0.7%	8 mg
VITAMIN C	40 mg	100% RDI#	20 mg

*BASED ON AN AVERAGE ADULT DIET OF 8700 KJ.
ALL VALUES CONSIDERED AVERAGES UNLESS OTHERWISE INDICATED

FIGURE 2 The nutritional information on a packet of biscuits (left) and a small carton of milk (right)

The nutritional information includes the quantities as a mass of each biomolecule per serving, the percentage of your daily intake (percentage of 8700 kJ) and the mass per 100 g of food. If the food item is a liquid, such as milk or soft drink, the values change to mass per volume quantity (see panel on the right-hand side of Figure 2).

Worked example 3.4F walks you through how to calculate energy from food products.

3.4F Worked example
Find me in your obook pro

3.4F Worked example
Video demonstration

FIGURE 3 Nutritional information panels are required, by law, on all food and drink products.

3.4 REAL-WORLD CHEMISTRY

What is a calorie?

Chemists and nutritionists have long argued over the term *calorie*. The calorie (cal), used by chemists, is the energy calculated using specific heat capacity. It is defined as the energy needed to raise the temperature of 1 g of water by 1°C at 1 atm. This is also nicknamed the *small calorie*.

Nutritionists recognise a calorie as a kilocalorie, the amount of heat required to raise the temperature of 1 kg of water by 1°C.

To distinguish between the two, food packaging uses a capital 'C' to refer to the dietary calorie, while chemists use the abbreviation *cal*.

Luckily for chemists, the unit has been changed to joules (J) and kilojoules (kJ), where 1 cal is equivalent to 4.2 J.

Apply your understanding

1 Construct a formula to convert a nutritionist's calorie to a chemical calorie.

2 The recommended calorie intake is 2500 for men and 2000 for women. Determine the number of kilojoules that this represents.

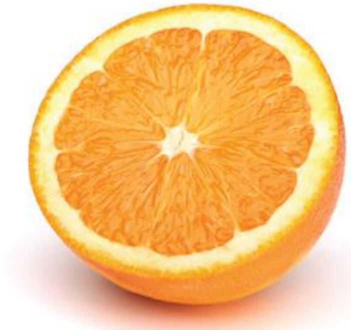

FIGURE 4 One medium-sized orange contains 68 cal or 286 J.

3.4 SKILL DRILL

Comparing experimental and theoretical heats of combustion

Key science skill: Analyse and evaluate data and investigation methods

A student measures the heat of combustion of ethanol in a spirit burner and records the following data in their results table:

Mass of the water (g)	150.0
Initial temperature of the water (°C)	16.2
Initial mass of the spirit burner (g)	253.8
Final temperature of the water (°C)	58.9
Final mass of the spirit burner (g)	252.1

Practise your skills

1 Using the mass of the ethanol, calculate the theoretical energy that should be transferred to the water.

2 Calculate the experimental energy transferred to the water.

3 Calculate the energy efficiency of the experiment.

4 Provide a reason for the difference between the theoretical and experimental results.

5 Identify whether this is a systematic or random error. Justify your answer.

6 Evaluate the validity of this experiment. Justify your answer.

Need help analysing and evaluating data and investigation methods? See Topic 1.8 (page 24).

3.4 CHECK YOUR LEARNING

Describe and explain

1 Explain why heat of combustion of fuels cannot be measured using a solution calorimeter.

2 Explain why the energy from impure fuels cannot be measured in kJ mol^{-1}.

3 Experimental data obtained from a spirit burner may not be accurate, but it may be precise. Explain why.

Apply, analyse and compare

4 A 1.25 g sample of butane (C_4H_{10}) was burnt using a spirit burner. The energy produced by the flame was used to heat a beaker filled with 400 g of water. The starting temperature of the water was 16.2°C and the final temperature was 38.5°C.

- **a** Calculate the experimental heat of combustion of butane in kJ mol^{-1}.
- **b** Calculate the percentage of energy lost in the process, based on the theoretical heat of combustion value of butane in your data book.

5 A 1.68 g sample of diesel was burnt using a spirit burner. The energy produced by the flame was used to heat a beaker filled with 750 g of water. The starting temperature of the water was 14.9°C and the final temperature was 36.4°C.

- **a** Calculate the enthalpy of the diesel (kJ g^{-1}).
- **b** Calculate the percentage of energy lost in the process based on the theoretical enthalpy value of diesel in your data book.

6 A 200 g block of parmesan cheese, with a serving size of 25 g, contains:

27.2 g/100 g fat
0.6 g/100 g carbohydrates
32.7 g/100 g protein.

For each serving, calculate:

- **a** the mass of each biomolecule
- **b** the energy of each biomolecule
- **c** the total energy
- **d** the energy as a percentage of the daily energy intake (8700 kJ).

7 A 1.00 L bottle of light milk, with a serving size of 250 mL, contains:

1.0 g/100 mL fat
4.6 g/100 mL carbohydrates
3.5 g/100 mL protein.

For each serving, calculate:

- **a** the mass of each biomolecule
- **b** the energy of each biomolecule
- **c** the total energy
- **d** the energy as a percentage of the daily energy intake (8700 kJ).

Design and discuss

8 Discuss the systematic errors and limitations of the spirit burner experiment when determining the heat of combustion of fuels.

9 Discuss the difference between drinking a cup of milk or Red Bull in the morning before school. Identify which provides you with more energy.

	Milk (250 mL)	**Red Bull (250 mL)**
Protein (g)	8.8	0
Fats (g)	8.8	0
Carbohydrates (g)	15.8	27.5

Assuming that carbohydrates are digested and absorbed into the bloodstream at a faster rate than fats and proteins, explain which drink would provide longer energy release.

CHAPTER 3 Review

Chapter summary

3.1

- Stoichiometry can be used to determine the mass or volume of a chemical species. This can be applied to determine the mass of major greenhouse gases generated from combustion reactions.
- Heat of combustion is used to calculate the heat energy released from combustion.

3.2

- Specific heat capacity is a measure of the amount of heat energy (J) that it takes to increase the temperature of 1 g of water by 1°C.
- The specific heat capacity of water, under ideal conditions, is $4.18\,\text{J}\,\text{g}^{-1}\,\text{K}^{-1}$ or $4.18\,\text{kJ}\,\text{kg}^{-1}$.
- The change in temperature of water is a good indicator of energy released from combustion reactions.

3.3

- Calorimetry is the technique of measuring the change in temperature of water caused by the heat energy generated from a reaction.
- Solution calorimeters measure the energy needed to increase the temperature of a solution by 1°C. Each calorimeter requires a different amount of energy and so each one must be calibrated to determine its calibration factor in $\text{J}\,°\text{C}^{-1}$.
- Once calibrated, reactions that produce aqueous solutions are conducted in the calorimeter. The temperature change of the solution can be used to calculate the amount of energy, in joules, that was absorbed or produced in the formation of products.

3.4

- Combustion reactions cannot be contained within a solution calorimeter and so a liquid fuel is placed in a pre-weighed spirit burner that is used to heat a known mass of water. The temperature increase is measured, and the mass of the spirit burner is re-weighed to determine the mass of fuel used to generate the temperature rise in the water.
- The energy content of various serving sizes of food can be calculated as the number of kilojoules per gram of each biomolecule within the food. These can also be calculated as a percentage of an adult's recommended daily dietary intake (8700 kJ).

Key formulas

Molar mass	$n = \frac{m}{M}$
Molar volume (at SLC)	$n = \frac{V}{V_m} = \frac{V}{24.8}$
Density	$d = \frac{\text{mass}}{\text{volume}}$
Energy from heat of combustion	$q = \Delta H°_{C} \times n$ or $q = \Delta H°_{C} \times m$
Specific heat capacity	$q = m \times c \times \Delta T$
Electrical energy in volts	$E = V \times I \times t$
Calibration factor	$\text{CF} = \frac{E}{\Delta T}$
Heat (or enthalpy) of solution	$\Delta H = \frac{E}{n}$ or $\Delta H = \frac{E}{m}$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as ‘Confidently’, ‘Mostly’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Mostly	Not really	Revision link
• use mass–mass, mass–volume and volume–volume stoichiometry in reactions involving the combustion of fuels, including to determine heat energy released, reactant and product amounts and net volume or mass of major greenhouse gases	☐	☐	☐	Go back to Topic 3.1
• use specific heat capacity of water to approximate the quantity of heat energy released during the combustion of a known mass of fuel and food	☐	☐	☐	Go back to Topic 3.2
• describe the properties of solution calorimetry, including determination of the calibration factor and consideration of the effects of heat loss	☐	☐	☐	Go back to Topic 3.3
• analyse temperature–time graphs obtained from solution calorimetry	☐	☐	☐	Go back to Topic 3.3
• calculate the energy transformation efficiency during the combustion of fuels and food as a percentage of chemical energy converted to useful energy	☐	☐	☐	Go back to Topic 3.4
• compare and calculate the energy values of foods containing carbohydrates, proteins and fats and oils	☐	☐	☐	Go back to Topic 3.4

Revision questions

Multiple choice

1 Which of the following fuels is the most sustainable?

A Ethanol

B Octane

C Diesel

D Kerosene

2 4.20 g of butane ($\Delta H = -2880\ \text{kJ mol}^{-1}$, $M = 58\ \text{g mol}^{-1}$) underwent complete combustion to heat an unknown mass of water. If the temperature of the water increased by 62°C, the mass of the heated water was:

A 1.24 g

B 0.805 g

C 9.7×10^{-5} g

D 805 g

Questions 3 and 4 refer to the following data.

Nitroglycerin undergoes explosive decomposition according to the following reaction:

$$4C_3H_5N_3O_9(l) \rightarrow 12CO_2(g) + 10H_2O(l) + 6N_2(g) + O_2(g) \qquad \Delta H = -5678\ \text{kJ mol}^{-1}$$

3 If 10 kg of nitroglycerin reacts, the mass of greenhouse gases produced at SLC is closest to:

A 7.80 kg

B 0.96 kg

C 5.82 kg

D 7.80 g

4 The energy (kJ) produced from 10 kg of nitroglycerin is:

A 62 335

B 62.3

C 15 584

D 250 116

Questions 5 and 6 refer to the following data.

A student experimentally determined the molar heat of combustion of ethanol ($M = 46.0\ \text{g mol}^{-1}$) using the equipment shown in Figure 1 (next page).

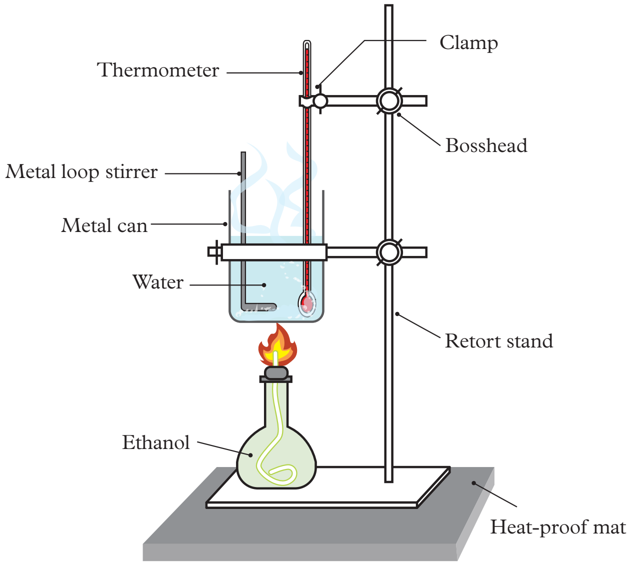

FIGURE 1 A spirit burner records the energy change in a combustion reaction.

The following data was recorded:

Mass of the water in the metal can = 80.0 g

Initial mass of the spirit burner = 125.40 g

Initial temperature of the water = 17.8°C

Final mass of the spirit burner = 124.55 g

Final temperature of the water = 38.9°C

5 The molar heat of combustion of ethanol is:

A 381 845.5 kJ mol^{-1}
B 704.0 kJ mol^{-1}
C 381.8 kJ mol^{-1}
D 8.30 kJ mol^{-1}

6 The spirit burner has an energy efficiency below 100% because:

A heat energy is absorbed from the surrounding environment.
B heat energy is lost to the surroundings.
C the water loses thermal energy as steam.
D the wick of the spirit burner is too close to the base of the metal can.

7 The complete combustion of methanol occurs according to the following equation:

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l) \qquad \Delta H = -1452\ \text{kJ mol}^{-1}$$

This means that:

A completely combusting 1.0 g of liquid methanol produces 1452 kJ of energy.
B 2.0 mol of liquid methanol produces 726 kJ of energy.
C 1.0 mol of liquid methanol produces 1452 kJ of energy.
D completely combusting 1.0 g of liquid methanol produces 22.69 kJ of energy.

8 Sodium hydroxide and nitric acid react according to the following equation:

$$NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

A 60 mL solution containing 0.5 M of NaOH is reacted in a calorimeter with 40 mL of 0.75 M HNO_3. The calorimeter has been calibrated and found to have a calibration factor of 14 J $°C^{-1}$; the temperature of the reaction increased by 8.9°C. The heat of neutralisation (in kJ mol^{-1}) is:

A 4153
B 2077
C 4.15
D 2.08

9 A 500 g block of tasty cheese contains 34.7 g/100 g fat, 0.8 g/100 g carbohydrates and 24.8 g/100 g protein. What percentage of an adult's dietary intake (8700 kJ) does 50 g of tasty cheese constitute?

A 19.75%
B 4.94%
C 9.74%
D 9.88%

10 A calorimeter is calibrated by applying a voltage of 5.6 V for 3.00 minutes. The current recorded during the calibration was 2.6 A and the temperature increase of the water was 0.875°C.

The calibration factor of the calorimeter is:

A 2995.2 J $°C^{-1}$
B 2995.2 kJ $°C^{-1}$
C 49.92 J $°C^{-1}$
D 49.92 kJ $°C^{-1}$

Short answer

Describe and explain

11 Define the term *specific heat capacity*.

12 Explain whether a molecule with stronger intermolecular interactions than water would have a higher or lower specific heat capacity.

13 Explain whether the calibration factor calculated in an uninsulated calorimeter would be higher or lower than in an insulated calorimeter.

14 Identify the most likely reasons moisture-free wood has a higher enthalpy than regular firewood.

Apply, analyse and compare

15 Compare a solution calorimeter with a spirit burner.

16 Calculate the molar heat of solution, in kJ mol^{-1}, of magnesium nitrate ($Mg(NO_3)_2$) when a 0.72 g sample of the salt is added to 123 g of water in a calorimeter and a temperature rise of 53°C is measured.

17 A 10.9 g sample of ethanol (C_2H_6O) was burnt in a spirit burner to heat a beaker filled with 320 g of water. The starting temperature of the water was 277 K and the final temperature was 72°C. Calculate the heat of combustion of the fuel in kJ mol^{-1}.

18 Calculate the energy released when 2.50 kg of propane is burnt in excess oxygen at SLC.

19 A 0.73 g sample of propan-1-ol (C_3H_8O) was burnt using a spirit burner. The energy produced by the flame was used to heat a beaker filled with 120 g of water. The starting temperature of the water was 17.6°C and the final temperature was 42.9°C.

a Calculate the experimental heat of combustion of propan-1-ol in kJ mol^{-1}.

b Calculate the energy efficiency of the process if propan-1-ol has a theoretical heat of combustion value of 2021 kJ mol^{-1}.

20 A 425 g can of tuna in springwater, with a serving size of 69 g, contains 1.0 g/100 g fat, 0.5 g/100 g carbohydrates, and 25.7 g/100 g protein.

For one serving, calculate:

a the mass of each biomolecule

b the energy of each biomolecule

c the total energy

d the energy as a percentage of the daily energy intake (8700 kJ).

21 A 2.00 L bottle of full-cream milk contains 3.4 g/100 mL fat, 4.4 g/100 mL carbohydrates, and 3.4 g/100 mL protein.

For one serving of 250 mL, calculate:

a the mass of each biomolecule

b the energy of each biomolecule

c the total energy

d the energy as a percentage of the daily energy intake (8700 kJ).

Design and discuss

22 Discuss how you might determine whether incomplete combustion has occurred using a spirit burner. Identify how you could modify the method to avoid incomplete combustion.

23 Discuss the difference in enthalpy values between ethane and ethanol from your data book and account for this difference.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER

4

Primary galvanic cells and fuel cells

KEY KNOWLEDGE

- → redox reactions as simultaneous oxidation and reduction processes, and the use of oxidation numbers to show the transfer of electrons from a reducing agent, oxidising agent and conjugate redox pairs
- → the writing of balanced half-equations (including states) for oxidation and reduction reactions, and the overall redox cell reaction in both acidic and basic conditions
- → the common design features and general operating principles of non-rechargeable (primary) galvanic cells converting chemical energy into electrical energy, including electrode polarities and the role of the electrodes (inert and reactive) and electrolyte solutions
- → the use and limitations of the electrochemical series in designing galvanic cells and as a tool for predicting the products of redox reactions, for deducing overall equations from redox half-equations and for determining maximum cell voltage under standard conditions
- → the common design features and general operating principles of fuel cells, including the use of porous electrodes for gaseous reactants to increase cell efficiency
- → the application of Faraday's Laws and stoichiometry to determine the quantity of galvanic or fuel cell reactant and product, and the current or time required to either use a particular quantity of reactant or produce a particular quantity of product
- → contemporary responses to challenges and the role of innovation in the design of fuel cells to meet society's energy needs, with reference to green chemistry principles: design for energy efficiency; and use of renewable feedstocks

Source: *VCE Chemistry Study Design (2023–2027)* reproduced by permission © VCAA

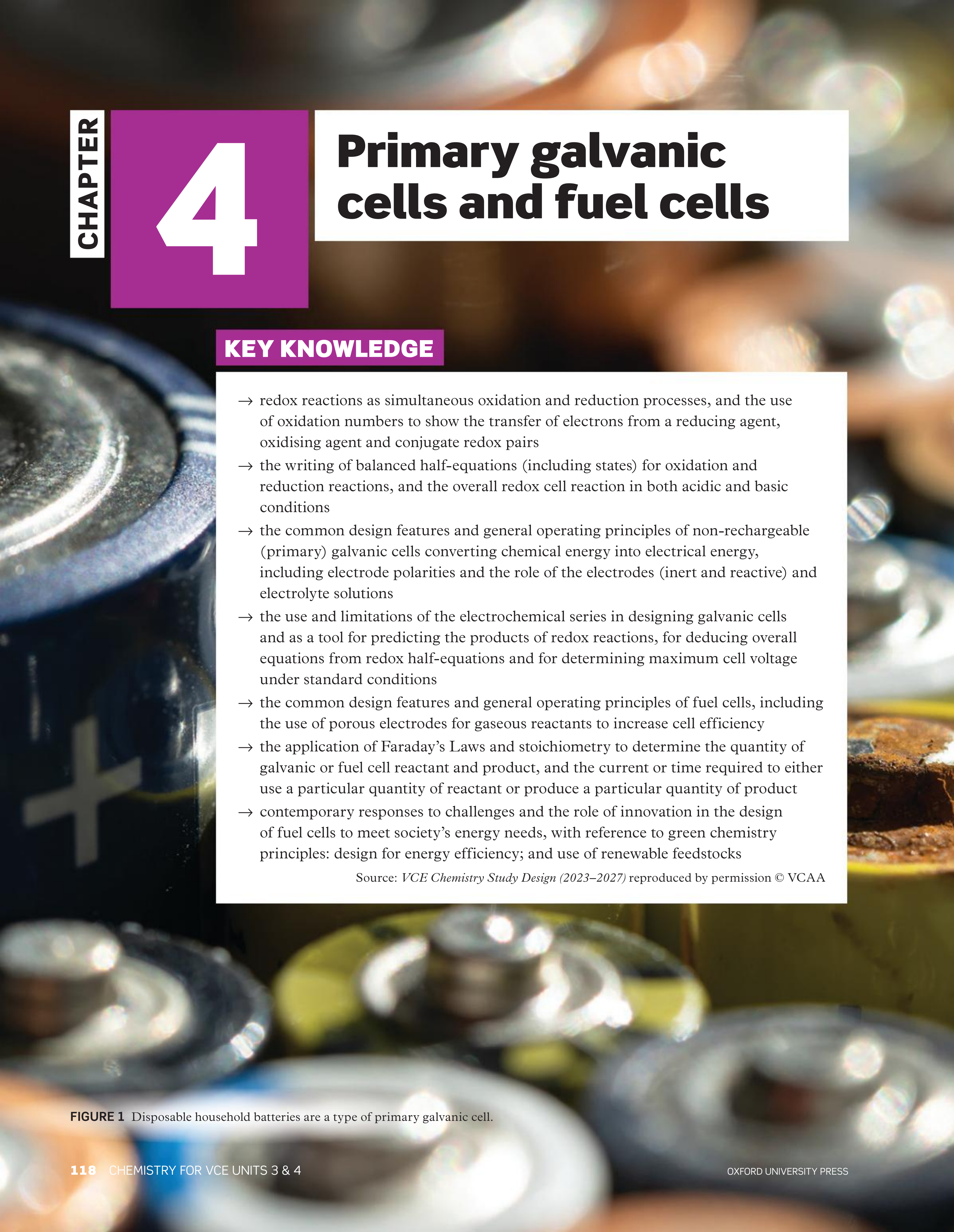

FIGURE 1 Disposable household batteries are a type of primary galvanic cell.

GROUNDWORK

In Chapter 4, you will learn about redox reactions and their applications in industry and everyday life, including primary galvanic cells and fuel cells as sources of energy.

This chapter will build on concepts you have already learnt in Chapter 2 and Unit 2. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

4A Explain what a redox reaction is.

4A Groundwork resource
Redox reactions

4B Explain what oxidising and reducing agents are and how you would be able to identify them.

4B Groundwork resource
Oxidising and reducing agents

4C Describe how biomass can be used to produce fuels.

4C Groundwork resource
Biofuels

PRACTICALS

4.4A **PRACTICAL: CONTROLLED EXPERIMENT**	How can we verify the electrochemical series using galvanic cells?	pro
4.4B **PRACTICAL: SIMULATIION**	Can you accurately predict cell voltage in a galvanic cell?	Page 505
4.7 **PRACTICAL: LITERATURE REVIEW**	Why are hydrogen fuel cells not widely used?	pro

4.1

Redox reactions

KEY IDEAS

In this topic, you will learn that:

- a redox reaction involves the transfer of electrons from one species to another
- oxidation (loss of electrons) and reduction (gain of electrons) occur simultaneously
- oxidation numbers are used to identify the oxidising agent and the reducing agent
- a conjugate redox pair consists of the oxidising agent and its conjugate reducing agent, or the reducing agent and its conjugate oxidising agent.

redox
a chemical reaction involving the transfer of electrons from one reactant to another

reduction
a process in which electrons are gained

oxidation
a process in which electrons are lost

Study tip

The abbreviation **OIL RIG** is useful for remembering the movement of electrons in oxidation and reduction. It stands for '**O**xidation **I**s **L**oss' and '**R**eduction **I**s **G**ain'.

In Unit 2, you were introduced to **redox** reactions. You may remember that they involve the transfer of negatively charged valence electrons between chemical species. This is like a financial transaction – when a substance loses one or more electrons, another substance accepts them.

Redox reactions

Redox is an abbreviation for two electron-transfer reactions that occur simultaneously. These reactions are called **reduction** and **oxidation**.

- Oxidation occurs when a chemical species loses one or more electrons.
- Reduction occurs when a chemical species gains one or more electrons.

As one reactant loses one or more electrons, the second reactant gains them to form new products. In Figure 1, species A has lost an electron and has therefore undergone oxidation. Species B has gained an electron, so it has undergone reduction. The redox reaction forms the ionic compound AB.

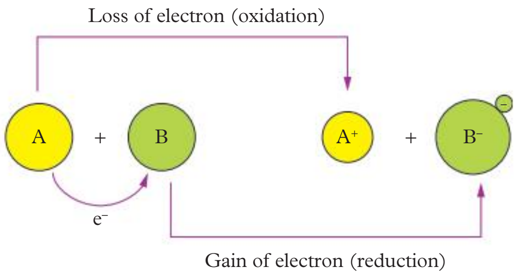

FIGURE 1 An electron is transferred from A to B to form a positive cation (A^+) and negative anion (B^-).

oxidation number
the degree of oxidation of a chemical species that indicates the number of electrons gained or lost by the species to bond in a compound; also known as oxidation state

Oxidation numbers

When an atom gains one or more electrons, it is reduced and gains a negative charge; its oxidation number decreases. On the other hand, a loss of electrons is oxidation; the atom gains a positive charge and its oxidation number increases. This is summarised in Table 1.

TABLE 1 Changes to atoms during redox reactions

	Reduction	Oxidation
Electrons are …	gained	lost
Atoms gain a … charge	negative	positive
Oxidation number …	decreases	increases

We can identify the number of electrons transferred between the reducing agent and the oxidising agent using **oxidation numbers** (or oxidation state). They help determine the total number of electrons that have been gained or lost to form a new product. They do not always represent the charge of individual chemical elements but are used as an indicator of how many electrons an atom has.

Oxidation numbers can be calculated for neutral elements, charged ions and covalent molecules. Table 2 outlines the rules for assigning oxidation numbers to different chemical species. Revisit how to determine oxidation numbers using Worked example 4.1A.

4.1A Worked example
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4.1A Worked example
Video demonstration

TABLE 2 Rules for assigning oxidation numbers

Rule	Examples
Elements and elemental molecules have an oxidation number of 0.	O_2, F_2, He, Fe, Zn, Li
Certain elements have common oxidation numbers when they are present in compounds.	• Group 1 metals (Li^+, Na^+, K^+) are always +1. • Group 2 metals (Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}) are always +2. • Hydrogen is +1 (except in metal hydrides, such as LiH, where it is –1). • Oxygen is –2 (except in peroxide, H_2O_2, where it is –1).
For monoatomic ions, the oxidation number is given by the charge on the ion.	Cu^{2+} has an oxidation number of +2. Na^+ has an oxidation number of +1.
In polyatomic ions, the sum of the oxidation numbers is equal to the charge of the ion.	In PO_4^{3-}, the sum of the individual oxidation numbers is –3, so P + (4 × O) = –3. Oxygen is –2 and phosphorus is +5. Substituting the numbers into the equation: 5 + (4 × –2) = –3.
In a neutral compound, the sum of oxidation numbers is equal to 0.	HCl is a neutral compound and has an oxidation number of 0. Because hydrogen is +1, chlorine must be –1.
The most electronegative element has a negative oxidation number.	NO_2 is a neutral compound. Oxygen is more electronegative and therefore has a negative oxidation number. Oxygen is –2 and 2 × –2 = –4. The sum of the oxidation numbers is 0, so nitrogen has an oxidation number of +4.

Oxidation numbers of transition metals

Transition metals, such as iron, have multiple oxidation states. This is also true of some non-metals, such as sulfur. These oxidation numbers are represented as roman numerals in brackets after the atom. They are also used in the name of ionic compounds. Examples include:

- copper, which has an oxidation number of +1 or +2 (rarely +3). Therefore, it can form Cu(I) or Cu(II). When naming copper in copper sulfate, it is either copper(I) sulfate or copper(II) sulfate.
- iron, which has an oxidation number of +2 or +3. Therefore, it is named Fe(II) or Fe(III). When naming iron in iron nitrate, it is either iron(II) nitrate or iron(III) nitrate.

If there is no rule for an atom, then its oxidation number must be determined using other atoms in the chemical species, as shown in Worked example 4.1B.

Identifying redox reactions

For a chemical reaction to be a redox reaction, both oxidation and reduction must be occurring. This means that one chemical species must be losing one or more electrons, while another gains one or more electrons. Redox reactions can only occur when an **oxidising agent** and **reducing agent** are present.

An oxidising agent:

- is responsible for causing oxidation
- is itself reduced in the redox reaction (gains electron(s)).

A reducing agent:

- is responsible for causing reduction
- is itself oxidised in the redox reaction (loses electron(s)).

Study tip

Charges are written with the symbol after the number; e.g., Al^{3+}. Oxidation numbers are written with the + or – symbol before the number; e.g. Al^{3+} has an oxidation number of +3.

4.1B Worked example
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4.1B Worked example
Video demonstration

oxidising agent
a reactant species that causes another to lose electrons and be oxidised; is itself reduced

reducing agent
a reactant species that causes another to gain electrons and be reduced; is itself oxidised

4.1 Real-world chemistry
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To identify a redox reaction, you can use oxidation numbers to help determine whether electrons have been gained and lost. Once you are confident with assigning oxidation numbers and identifying the redox half-equations, apply your understanding to corrosion of the Statue of Liberty in Real-world chemistry 4.1.

Identifying conjugate redox pairs

conjugate redox pair
two chemical species that differ by the number of electrons that are transferred in the redox reaction

Each oxidising and reducing agent involved in a redox reaction will have its own **conjugate redox pair**. This consists of the oxidising agent (reactant) and its conjugate reducing agent (product), or the reducing agent (reactant) and its conjugate oxidising agent (product).

The two species in a conjugate redox pair sit on opposite sides of the reaction arrow. They differ by one or more electrons, depending on how many electrons are transferred in the redox reaction.

Study tip

Two species in a conjugate redox pair differ by the number of electrons transferred in the redox reaction.

In Figure 2, species A is the reducing agent because it has lost an electron. Its conjugate oxidising agent is the A^+ cation, which has the ability to gain an electron (undergo reduction). Species B is the oxidising agent because it has gained an electron. Its conjugate reducing agent is the B^- anion, which has the ability to lose an electron (undergo oxidation). The conjugate redox pairs in this reaction are A/A^+ and B/B^-.

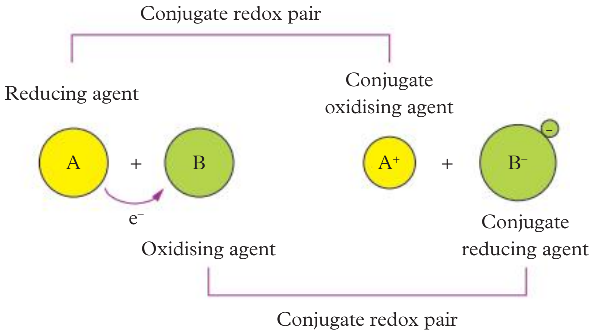

FIGURE 2 Conjugate redox pairs differ by one electron.

Worked example 4.1C walks you through how to identify conjugate redox pairs.

Study tip

If you are asked to identify a redox reaction in an exam question, consider the oxidation numbers of hydrogen and oxygen first. Hydrogen is always +1, except in metal hydrides (e.g. NaH), where it is –1. Oxygen is always –2, except in peroxide (H_2O_2), where it is –1.

FIGURE 3 The Fe^{2+} and Fe^{3+} ions are conjugate redox pairs because they differ by one electron. They also differ in colour.

4.1C WORKED EXAMPLE

IDENTIFYING CONJUGATE REDOX PAIRS

Consider the following chemical equation:

$$H_2O(l) + SO_2(g) + I_2(g) \rightarrow H_2SO_4(aq) + 2HI(aq)$$

Identify the oxidising and reducing agents, by assigning oxidation numbers to all atoms according to the rules in Table 2. Then, identify the conjugate redox pairs.

Solution

Think	Do	
Step 1: Determine the sum of all oxidation numbers.	Reactants side: H_2O is a neutral compound, so the sum of its oxidation numbers is zero.	Products side: H_2SO_4 is a neutral compound, so the sum of its oxidation numbers is zero.
	SO_2 is a neutral compound, so the sum of its oxidation numbers is zero.	HI is a neutral compound, so the sum of its oxidation numbers is zero.
	I_2 is a molecular element, so it has an oxidation number of zero.	
Step 2: Using the number of atoms that make up the compound, write an equation to calculate the oxidation numbers.	$(2 \times H) + (O) = 0$	
	$(S) + (2 \times O) = 0$	$(2 \times H) + (S) + (4 \times O) = 0$
	$(2 \times I) = 0$	$(H) + (I) = 0$
Step 3: Find the missing oxidation numbers. Note: Oxygen is always –2 (unless in hydrogen peroxide) and hydrogen is always +1 unless in a metal hydride.	$(2 \times +1) + (-2) = 0$ $H = +1$ and $O = -2$	
	$(S) + (2 \times -2) = 0$ $(S) + (-4) = 0$ $S = +4$	$(2 \times +1) + (S) + (4 \times -2) = 0$ $(2) + (S) + (-8) = 0$ $S = +6$
	$(2 \times I) = 0$ $I = 0$	$(+1) + (I) = 0$ $I = -1$
Step 4: Write the oxidation numbers on the top of each atom in the chemical equation to help identify the oxidising agent and reducing agent.	$\overset{+1-2}{H_2O}(l) + \overset{+4-2}{SO_2}(g) + \overset{0}{I_2}(aq) \rightarrow \overset{+1+6-2}{H_2SO_4}(aq) + 2\overset{+1-1}{HI}(aq)$ Iodine undergoes reduction, as its oxidation number decreases from 0 to –1. It causes another reactant to lose electrons, and is therefore the oxidising agent. Sulfur undergoes oxidation, as its oxidation number increases from +4 to +6. It causes another reactant to gain electrons, and is therefore the reducing agent.	
Step 5: Identify the conjugate redox pairs.	The conjugate redox pairs are I_2/I^- and SO_2/SO_4^{2-}.	

4.1 CHECK YOUR LEARNING

Describe and explain

1 Define *oxidation number.*

2 Explain what happens to oxidation numbers in reduction reactions.

3 Describe the processes of reduction and oxidation, and explain how they are related.

Apply, analyse and compare

4 Determine the oxidation numbers of the atoms in the following chemical substances.

- **a** O_2
- **b** NO_2
- **c** SO_4^{2-}
- **d** CH_3COO^-
- **e** NaOH
- **f** H_2O_2
- **g** NaH
- **h** PO_4^{3-}

5 Determine the oxidising and reducing agents in the following chemical equations. Use their oxidation numbers to justify your answers.

- **a** $2Fe(OH)_3(aq) + 3OCl^-(aq) \rightarrow 2FeO_4^{2-}(aq) + 3Cl^-(aq) + H_2O(l) + 4H^+(aq)$
- **b** $Fe^{2+}(aq) + 6H^+(aq) + VO_4^{3-}(aq) \rightarrow Fe^{3+}(aq) + VO^{2+}(aq) + 3H_2O(l)$
- **c** $Cr_2O_7^{2-}(aq) + 16H^+(aq) + C_2H_5OH(l) \rightarrow 4Cr^{3+}(aq) + 2CO_2(g) + 11H_2O(l)$

6 Identify the conjugate redox pairs in Question 5.

4.2 Half-equations and overall equations

KEY IDEAS

In this topic, you will learn that:

- oxidation and reduction reactions can be expressed using simple half-equations
- oxidation and reduction reactions in acidic and basic conditions can be expressed using complex half-equations
- oxidation and reduction half-equations can be combined to develop an overall redox equation.

Redox equations can become complicated due to the number of chemical species involved. For this reason, they are balanced in **half-equations**, where the oxidation equation is represented as one half-equation and the reduction equation is represented as the second half-equation. These equations are then combined to form an **overall redox equation**.

These equations show the chemical species that participate in the redox reaction. All other species act as **spectator ions**, which do not participate in electron transfer and are therefore not included.

half-equation
an equation which represents half of a redox equation and is either an oxidation or reduction half-equation; it shows electron transfer

overall equation
an overall redox reaction that combines the two half-equations after electrons have been balanced and cancelled out

spectator ion
an ion that exists in the same form on the reactant side and product side of a redox reaction; there is no change in its oxidation number or state

Half-equations

Every redox reaction is made up of two half-reactions. The oxidation half-equation shows an atom losing electrons, while the reduction half-equation shows an atom gaining electrons. To properly balance overall redox equations (which you will look at later), it is essential to represent the electrons in the half-equations.

Simple half-equations

Simple half-equations are straightforward to balance. Remember from Units 1 and 2 that each half-equation will show the conjugate redox pair and the electrons being transferred. Let's look at the reaction between copper(II) sulfate and zinc metal below, with oxidation numbers assigned. In this case, it is easier to deal with sulfate when its oxidation number is equal to its charge of –2.

$$\overset{+2\,+2-2}{CuSO_4(aq)} + \overset{0}{Zn(s)} \rightarrow \overset{0}{Cu(s)} + \overset{+2\,+2-2}{ZnSO_4(aq)}$$

Copper is reduced because its oxidation number has decreased. $Cu^{2+}(aq)$ has gained two electrons to become Cu(s). The electrons are shown on the left-hand side of the reaction arrow to balance the 2+ charge. Therefore, the half-equation is:

$$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$$

Zinc is oxidised because its oxidation number has increased. Zn(s) has lost two electrons to become $Zn^{2+}(aq)$. The electrons are shown on the right side of the reaction arrow to balance the 2+ charge. Therefore, the half-equation is:

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$$

Sulfate is a spectator ion since its oxidation number and state are the same on both sides of the equation. Therefore, it is excluded from the half-equations.

Study tip

Draw a line down the middle of the word *redox*: red | ox. 'Red' is on the left-hand side, so for reduction reactions, the electrons appear on the left side of the reaction arrow. 'Ox' is on the right-hand side, so for oxidation, they appear on the right side of the arrow.

Complex half-equations

Often, redox half-equations cannot be balanced by simply adding electrons. In the case of more **complex half-equations**, where hydrogen or oxygen is included, the balancing of equations follows the KOHES method. Complex redox reactions occur in acidic or basic conditions.

complex half-equation
a half-equation for an oxidation and reduction equation that occurs under acidic or basic conditions

In *acidic* conditions, the following steps are used (see Worked example 4.2A):

1 K – Balance the **key** elements (everything except oxygen and hydrogen).
2 O – Balance the **oxygen** by adding $H_2O(l)$ molecules to the opposite side. Each water molecule contains one oxygen atom.
3 H – Balance the **hydrogen** by adding $H^+(aq)$ ions to the opposite side.
4 E – Balance the charge by adding **electrons** to one side.
5 S – Assign **states** to all reactants and products except electrons, which have no state.

In *basic* (alkaline) conditions, add the following steps (see Worked example 4.2B):

6 Identify the number of H^+ ions and add the same number of **hydroxide** ($OH^-(aq)$) ions to both sides of the equation.
7 On the side that has both H^+ and OH^- ions, combine these to form **water** ($H_2O(l)$) molecules: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
8 **Cancel** out the water molecules so that only one side of the equation has them.

Note: The overall charge on each side of the reaction arrow remains the same.

Study tip

Remember to include states for all species in a redox equation (e.g. (g), (l), (aq)) – except for electrons, which do not have a state.

4.2A WORKED EXAMPLE

WRITING COMPLEX HALF-EQUATIONS IN ACIDIC CONDITIONS

Complete the half-equation for the permanganate ion (MnO_4^-) forming a manganese ion (Mn^{2+}) in acidic conditions and deduce whether this is an oxidation or reduction half-equation.

$$MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$$

Solution

Think	Do
Step 1: Balance the **key** elements.	There is one manganese atom on the left and one on the right. Key elements are already balanced. $MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$
Step 2: Balance the **oxygen** by adding H_2O molecules to the opposite side.	There are four oxygen atoms on the left, so we add four H_2O molecules to the right. $MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O$
Step 3: Balance the **hydrogen** by adding H^+ ions to the opposite side.	There are eight hydrogen atoms on the right, so we add eight H^+ ions to the left. $MnO_4^-(aq) + 8H^+ \rightarrow Mn^{2+}(aq) + 4H_2O$
Step 4: Balance the charge by adding **electrons** to one side.	The total charge on the left is –1 (for MnO_4^-) plus +8 (for $8H^+$) = +7. The total charge on the right is +2 (for Mn^{2+}) plus 0 (for H_2O) = +2. We need $5e^-$ on the left, since 7 – 5 = +2. $MnO_4^-(aq) + 8H^+ + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O$
Step 5: Add **states** to all reactants and products except electrons, which have no state.	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$
Step 6: Identify whether the reaction is oxidation or reduction.	Electrons are gained, so this is a reduction half-equation.

Study tip

You can look at either the change in oxidation number or the side of the half-equation that the electrons are on to determine whether it is an oxidation or reduction half-equation.

4.2B WORKED EXAMPLE

WRITING COMPLEX HALF-EQUATIONS IN BASIC CONDITIONS

Complete the half-equation for ethanol (CH_3CH_2OH) forming carbon dioxide (CO_2) in basic conditions and deduce whether this is an oxidation or reduction half-equation.

$$CH_3CH_2OH(aq) \rightarrow CO_2(g)$$

Solution

Think	Do
Step 1: Balance the **key** elements.	There are two carbon atoms on the left and one on the right. We need to add a coefficient to the right side. $CH_3CH_2OH(aq) \rightarrow 2CO_2(g)$
Step 2: Balance the **oxygen** by adding H_2O molecules to the opposite side.	There is one oxygen atom on the left and four on the right, so we add three H_2O molecules to the left. $CH_3CH_2OH(aq) + 3H_2O \rightarrow 2CO_2(g)$
Step 3: Balance the **hydrogen** by adding H^+ ions to the opposite side.	There are 12 hydrogen atoms on the left, so we add 12 H^+ ions to the left. $CH_3CH_2OH(aq) + 3H_2O \rightarrow 2CO_2(g) + 12H^+$
Step 4: Balance the charge by adding **electrons** to one side.	The total charge on the left is 0 (for CH_3CH_2OH) plus 0 (for $3H_2O$) = 0. The total charge on the right is 0 (for $2CO_2$) plus +12 (for $12H^+$) = +12. We need $12e^-$ on the right, since +12 – 12 = 0. $CH_3CH_2OH(aq) + 3H_2O \rightarrow 2CO_2(g) + 12H^+ + 12e^-$
Step 5: Add **states** to all reactants and products except electrons, which have no state.	$CH_3CH_2OH(aq) + 3H_2O(l) \rightarrow 2CO_2(g) + 12H^+(aq) + 12e^-$ Ethanol is a base, so we need to follow the next steps to balance the half-equation.
Step 6: Identify the number of H^+ ions and add the same number of **hydroxide** ions to both sides of the equation.	There are 12 H^+ ions, so we add 12 OH^- ions to both sides. $CH_3CH_2OH(aq) + 3H_2O(l) + 12OH^-(aq) \rightarrow$ $2CO_2(g) + 12H^+(aq) + 12e^- + 12OH^-(aq)$
Step 7: On the side which has both H^+ and OH^- ions, combine them to form **water** molecules.	There are 12 H^+ and OH^- ions on the right, so we combine them to make $12H_2O$. $CH_3CH_2OH(aq) + 3H_2O(l) + 12OH^-(aq) \rightarrow$ $2CO_2(g) + 12H_2O(l) + 7e^-$
Step 8: **Cancel** out the water molecules so that only one side of the equation has them.	There are three H_2O molecules on the left and $12H_2O$ molecules on the right, so we remove three from each side. $CH_3CH_2OH(aq) + 12OH^-(aq) \rightarrow 2CO_2(g) + 9H_2O(l) + 7e^-$
Step 9: Identify whether the reaction is oxidation or reduction.	Electrons are lost, so this is an oxidation half-equation.

Study tip

Be careful to check whether the redox reaction is occurring in acidic or basic conditions. In basic conditions, you will need to follow the extra steps.

Overall redox equations

When the oxidation and reduction half-equations have been identified, they can be combined to form an overall redox equation. To do this, both half-equations must have the same number of electrons so that they can cancel out in the final equation. You can see this in Worked example 4.2C.

FIGURE 1 Before combining the half-equations, the electrons must made equal and cancelled out.

4.2C Worked example
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4.2C Worked example
Video demonstration

4.2 CHALLENGE

Writing overall equations using complex half-equations

Butan-1-ol ($CH_3CH_2CH_2CH_2OH$) is oxidised to butanoic acid ($CH_3CH_2CH_2COOH$) in a solution of acidified sodium dichromate ($Na_2Cr_2O_7$) according to the following half-equations.

$$Cr_2O_7^{\ 2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

$$CH_3CH_2CH_2CH_2OH(aq) + H_2O(l) \rightarrow CH_3CH_2CH_2COOH(aq) + 4H^+(aq) + 4e^-$$

Combine the half-equations and write the balanced overall equation for the reaction.

4.2 CHECK YOUR LEARNING

Describe and explain

1 Describe what all oxidation half-equations have in common.

2 Describe what all reduction half-equations have in common.

3 Explain why overall redox equations must have the same number of electrons on the reactant and product sides.

Apply, analyse and compare

4 Write the half- and overall equations for the following reactions and balance the overall equations, remembering to identify the half-equations as reduction or oxidation.

- **a** $Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$
- **b** $S(s) + F_2(g) \rightarrow SF_6(g)$
- **c** $N_2(g) + Cl_2(g) \rightarrow N_2Cl_4(g)$
- **d** $CO_2(g) + H_2O(l) \rightarrow C_6H_{12}O_6(aq) + O_2(g)$

5 Identify the oxidising and reducing agents in each reaction from Question 4.

Design and discuss

6 Copper found in ancient burial sites has been dated as early as 9000 BC. Investigate how copper was smelted in ancient times in comparison to today. Communicate the key differences in the processes.

7 Research tarnished silverware and communicate a procedure for cleaning the silverware using redox chemistry. You must include chemical equations in your method.

4.3 Features of non-rechargeable galvanic cells

KEY IDEAS

In this topic, you will learn that:

- galvanic cells produce electrical energy from chemical energy
- single-use batteries are primary galvanic cells.

FIGURE 1 Chemical energy is converted to electrical energy in non-rechargeable batteries.

When negatively charged electrons are transferred between chemical species in a redox reaction, energy is released. With the correct set-up, the chemical energy can be converted into electrical energy. Electrical energy can also be converted into chemical energy. These processes occur within devices called **electrochemical cells**.

Electrochemical cells in which chemicals are used to generate electricity are called **galvanic cells**. You encounter galvanic cells every day, in the cells and batteries that power many of your electronic devices. In this topic, we will explore how non-rechargeable batteries work.

electrochemical cell
a device in which electrical energy can be generated from chemical energy, or chemical energy from electrical energy

galvanic cell
a type of electrochemical cell where the redox reaction is separated into reduction and oxidation half-cells; these are connected using a circuit to generate electricity

half-cell
the component of a galvanic cell that contains either the oxidation or reduction redox reaction

Galvanic cells

In a galvanic cell, a spontaneous (occurring naturally, without energy input) redox reaction occurs to produce electrical energy. The oxidation and reduction half-reactions occur in separate electrochemical cells, called **half-cells**. One cell produces electrons (the site of oxidation) and the other consumes them (the site of reduction). The cells are connected by wires through which a current can pass, like in an electric circuit.

Galvanic cells consist of several key features:

- half-cells – containing the redox reactants and products
- electrodes – metals that act as the anode and cathode
- an external circuit – consisting of wires (through which electrons move) and a voltmeter
- an internal circuit – consisting of a salt bridge (through which ions move)

Let’s look at each feature and their role in a galvanic cell in more detail.

FIGURE 2 The set-up of a laboratory-based galvanic cell

Half-cells

In a galvanic cell, two redox half-reactions are contained within separate half-cells. In laboratory-based galvanic cells, the half-cells consist of a beaker that contains an **electrolyte solution** and an **electrode**. The electrode can be an inert (unreactive) substance, such as carbon/graphite (C) or platinum (Pt), or a potentially reactive metal of the solution (e.g. a copper electrode in a copper(II) sulfate ($CuSO_4$) solution) (Figure 3). It conducts electricity into or out of a half-cell.

Half-cells are represented as the metal and the metal ion. In Figure 3, the metal of the solution is used, so the half-cell is $Cu(s)/Cu^{2+}(aq)$, the conjugate redox pair. If an **inert electrode** is used, the half-cell is represented as $C(s)/Cu^{2+}(aq)$ or $Pt(s)/Cu^{2+}(aq)$. The spectator ions do not react and are typically not represented in equations.

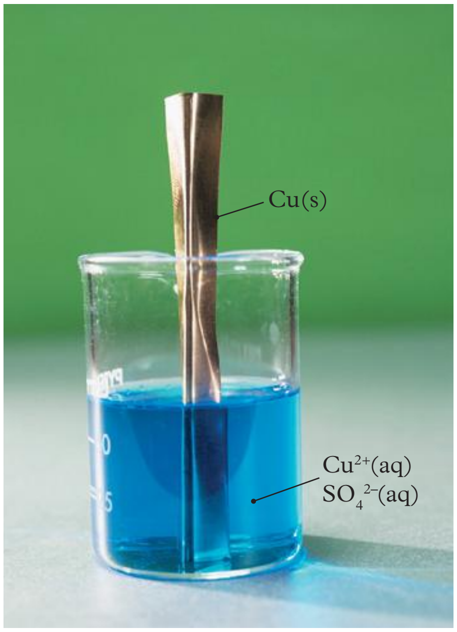


FIGURE 3 A copper half-cell containing copper metal ($Cu(s)$) and a solution of copper(II) sulfate ($Cu^{2+}(aq)$ and $SO_4^{2-}(aq)$). The $SO_4^{2-}(aq)$ is a spectator ion.

electrolyte solution
an aqueous solution containing ions

electrode
a strip or rod that conducts electricity; this is the surface on which the oxidation or reduction reactions occur

inert electrode
an electrode in a half-cell that does not participate in the reaction – i.e. it is neither a reactant nor product

Oxidation half-cell

When a galvanic cell is in operation, half-reactions occur in each half-cell. One reaction produces electrons. This occurs in the oxidation half-cell.

During the oxidation of a solid metal, the metal oxidises and breaks down in the solution to form a metal ion. The electrode must supply the reactant for the reaction (the reducing agent), so it needs to be a **reactive electrode** made of the metal itself. It cannot be made of an inert substance.

reactive electrode
in a half-cell, an electrode that participates in the reaction and is oxidised during the discharge reaction

However, you do not always need to use a reactive metal electrode. Some cells involve oxidation of gases or aqueous ions. One example is the oxidation of iron(III) to iron(II). Using an electrode made of solid iron will cause an additional reaction between iron(II) and the solid iron, so the electrode must be inert.

$$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$$

$$Fe^{2+}(aq) + 2\,e^- \rightarrow Fe(s)$$

In the oxidation half-cell, the cations become dissolved in the electrolyte solution. You will understand the role of the electrolyte solution better once you learn about the internal circuit.

Reduction half-cell

The other reaction consumes electrons. This occurs when the ions in the solution gain electrons to form a solid metal. This is the reduction half-cell.

The metal formed will plate (cover) the electrode because the electrode is the site of the reaction. Because the electrode will not break down in this half-cell, it can be made of an inert substance.

While the electrode often supplies the reducing agent in the oxidation half-cell, the electrolyte solution supplies the reactant (the oxidising agent) in the reduction half-cell. For example, the electrolyte solution must contain Cu^{2+} ions in a $Cu(s)/Cu^{2+}(aq)$ half-cell.

Study tip

Reactants are kept separate (in two half-cells) because when they are in direct contact, they react violently in an exothermic reaction and produce a large amount of heat energy, instead of electrical energy.

Electrodes

The electrodes that are used in the half-cells are called the **anode** and the **cathode**.

Oxidation occurs at the anode. **O**xidation **i**s the **l**oss of electrons (OIL) from a reactant. As electrons are generated at the surface of the anode, the anode has a negative charge. In an oxidation reaction, electrons are lost as the reactive metal anode breaks down to form positive metal cations. Over time, the concentration of metal ions in the solution increases.

anode
the negatively charged electrode

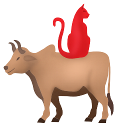
cathode
the positively charged electrode

Reduction occurs at the cathode. **R**eduction **i**s the **g**ain of electrons (RIG) by a reactant. Electrons are consumed at the cathode and the surface of the cathode is positively charged.

Charges at the anode and cathode are often referred to as **electrode polarity**.

Study tip

An easy way to remember that **red**uction occurs at the **cat**hode is to use the acronym **RED CAT**. The **an**ode is where **ox**idation occurs, so the acronym is **AN OX**. Galvanic cells involve AN OX and a RED CAT.



electrode polarity
the charge of an electrode

External circuit

Electrons are produced at the negative anode and then move to the positive cathode, where they are consumed. This generates a current.

The current flows through an **external circuit** which connects the electrodes. This consists of wires and a voltmeter (or multimeter). The voltmeter measures the voltage (in volts, V) generated by the galvanic cell when current flows from the anode to the cathode.

FIGURE 4 Connecting wires are very important components of the external circuit.

external circuit
the electric circuit (wires) which allow for electron movement from anode to cathode; also includes the voltmeter

Salt bridge

The **salt bridge** is the **internal circuit** that connects the two half-cells. It is saturated with another electrolyte solution. The purpose of the salt bridge is to supply ions to the half-cells in order to balance the charges formed during the redox reaction and to complete the circuit. Commonly used electrolyte solutions include KNO_3, NaCl, KOH and HCl.

The movement of ions in a galvanic cell is summarised in Table 1.

salt bridge
internal circuit connecting two half-cells; allows for the flow of charge by moving ions; cations move to the cathode and anions move to the anode

TABLE 1 Movement of ions in a galvanic cell

	Oxidation half-cell	Reduction half-cell
Electrons are …	produced	consumed
at the …	anode.	cathode.
Positive ions are …	formed as the metal electrode breaks down into solution.	removed from the solution and either plate (cover) the electrode or are converted to an uncharged compound.
This causes a build-up of …	positive metal ions.	negative spectator ions, which remain in the solution.
… ions migrate towards this half-cell to balance the charge.	Negative	Positive

Study tip

In the salt bridge, negative anions move to the negative anode and positive cations move to the positive cathode.

internal circuit
the salt bridge that allows ions to flow

Galvanic cells can be represented using a **cell diagram**, which shows all the features and the directions of electron and ion movement (Figure 5). Worked example 4.3 shows you how to draw your own galvanic cell.

cell diagram
a diagram of a galvanic cell that has all the key components, reaction mechanisms and electron/ion movement labelled

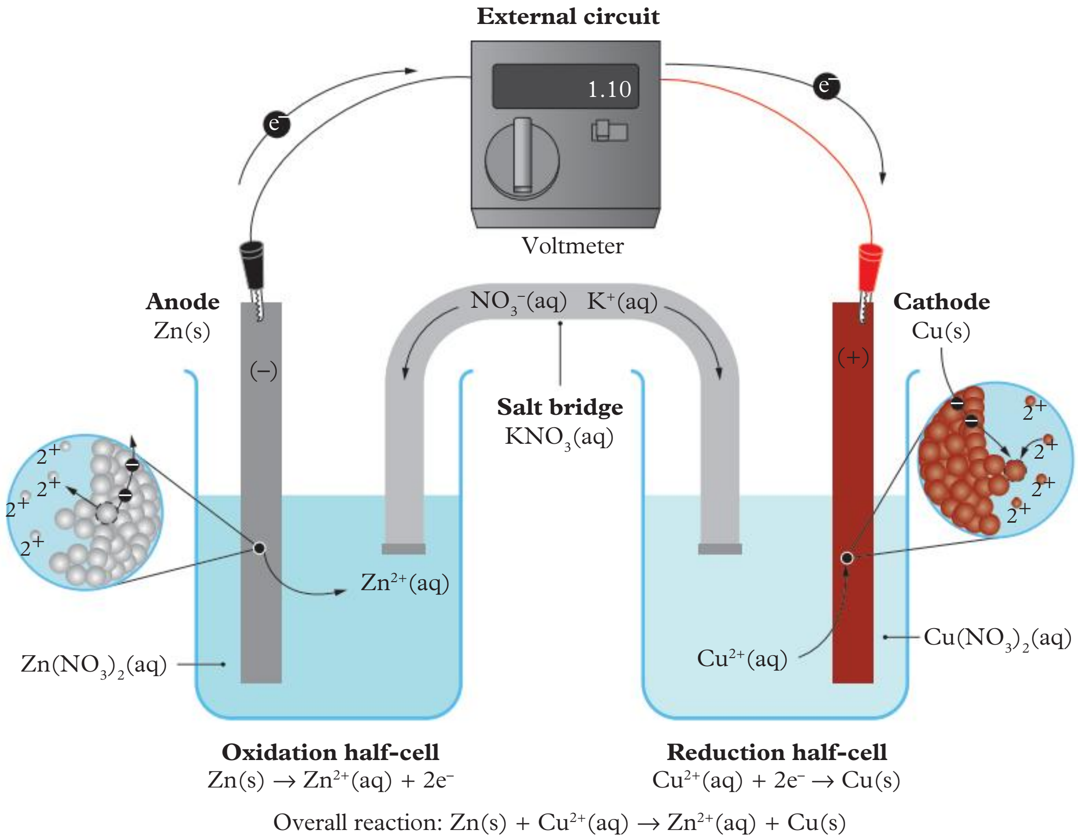


FIGURE 5 A galvanic cell for the reaction between solid zinc and copper ions

Study tip

Chemistry and Physics deal with the polarity of electrodes differently. If you study both, be careful not to get confused. In Chemistry, the galvanic cell has a positive cathode and a negative anode.

4.3 Worked example
Find me in your obook pro

4.3 Worked example
Video demonstration

Primary galvanic cells – non-rechargeable batteries

A galvanic cell is often called a **primary cell** or battery, and it can only be discharged (and not recharged).

Discharge is the process of converting chemical to electrical energy. As the solid anode is oxidised to form ions in the solution, it breaks down and no longer completes the circuit. Once the chemical reaction occurring in the cell has used all the reactants, the battery dies and must be thrown away.

The design of primary cells does not allow the reactions to be reversed. If the battery is rechargeable, it is called a secondary cell (which you will learn about in Chapter 7).

FIGURE 6 Primary cells must be disposed of after use.

primary cell
a non-rechargeable battery; a galvanic cell that produces an electrical current

discharge
the conversion of chemical to electrical energy

Study tip

Galvanic cells are sometimes drawn with a *load* in the circuit as well as the voltmeter/ammeter. This can be a light bulb, appliance or any kind of device that uses the electrical energy that is generated by the cell.

Non-rechargeable batteries do not have a salt bridge; instead, they contain electrolytes. Each electrode is in direct contact with an electrolyte, which is either acidic (supplies H^+ ions) or alkaline (supplies OH^- ions).

Primary cells do not produce high voltages. They are mostly used for low-current devices; however, they can be connected in series to produce larger voltages.

Alkaline batteries

Alkaline batteries are named after the electrolyte contained in them. They include some of the most commonly purchased batteries, such as AA and AAA batteries; they are also called zinc–carbon galvanic cells. A fun fact is that the As in AA and AAA do not actually stand for anything. They are used to denote the size of the battery.

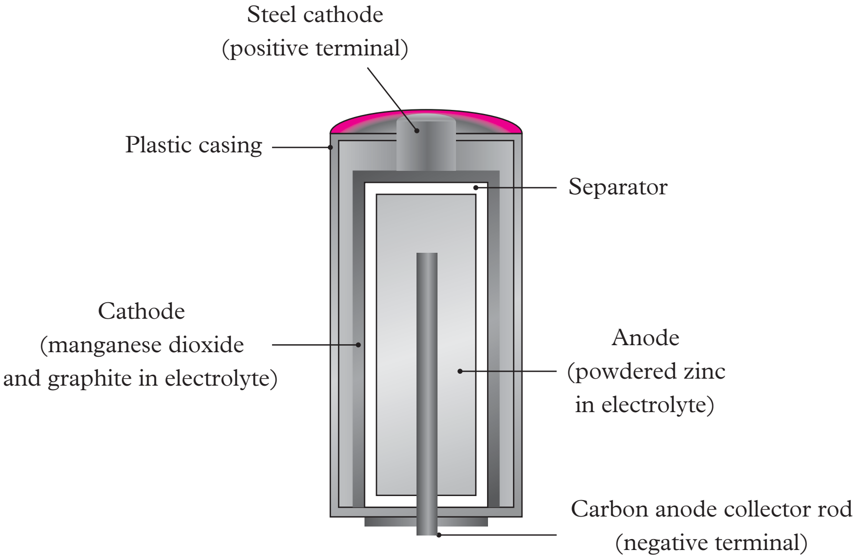


FIGURE 7 An alkaline battery

Study tip

The new study design does not require you to remember specific cells. You just need to understand how the reactions occur within them.

An alkaline cell (Figure 7) has a powdered Zn(s) anode in a potassium hydroxide (basic KOH) electrolyte. The cathode is made of a mixture of manganese dioxide ($MnO_2(s)$) and graphite. The separator makes sure that the reactants do not directly react spontaneously with one another and produce thermal energy instead.

These types of batteries produce approximately 1.5 V, which decreases with time as the reactants contained within the cell are consumed in the reaction.

At the anode, solid zinc is oxidised in the presence of hydroxide ions to form zinc hydroxide and two electrons:

$$Zn(s) + 2OH^-(aq) \rightarrow Zn(OH)_2(s) + 2e^-$$

At the cathode, manganese dioxide is reduced in the aqueous electrolyte. Two electrons are consumed to form manganese(III) oxide and hydroxide ions:

$$2MnO_2(s) + H_2O(l) + 2e^- \rightarrow Mn_2O_3(s) + 2OH^-(aq)$$

The overall reaction equation for the alkaline battery is:

$$Zn(s) + 2MnO_2(s) + H_2O(l) \rightarrow Zn(OH)_2(s) + Mn_2O_3(s)$$

FIGURE 8 9V, AA and AAA batteries

Button cells

Button cells are much smaller in size than alkaline cells. They can be used in small electrical devices, such as watches, hearing aids, calculators and cameras. Despite their size, they can produce similar voltages of approximately 1.5 V or sometimes greater.

One of the more efficient types of button cell, the silver oxide cell, consists of a powdered zinc anode with an alkaline potassium hydroxide (KOH) electrolyte. The cathode is made of a mixture of silver oxide and graphite (Figure 9).

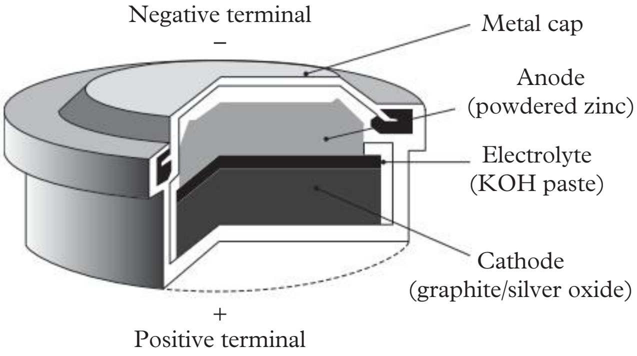


FIGURE 9 A silver oxide button cell

At the anode, zinc is oxidised in the presence of hydroxide ions from the electrolyte paste to form zinc oxide and two electrons:

$$Zn(s) + 2OH^-(aq) \rightarrow Zn(OH)_2(s) + 2e^-$$

At the cathode, silver oxide is reduced in the aqueous electrolyte. Two electrons are consumed to form silver metal and hydroxide ions:

$$Ag_2O(s) + H_2O(l) + 2e^- \rightarrow 2Ag(s) + 2OH^-(aq)$$

The overall reaction equation for this type of button cell is:

$$Zn(s) + Ag_2O(s) + H_2O(l) \rightarrow 2Ag(s) + Zn(OH)_2(s)$$

A button cell is more expensive than an alkaline battery, but it can produce a stable voltage over long periods of time.

FIGURE 10 Button batteries

4.3 CHECK YOUR LEARNING

Describe and explain

1 Explain how a galvanic cell is able to generate electricity. Reference ion and electron movement in your answer.

2 Describe the purpose of a salt bridge in a galvanic cell.

Apply, analyse and compare

3 Consider the following half-equations:

$$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$$
$$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$$

- **a** Identify the oxidising agent and reducing agent.
- **b** The oxidising and reducing agents identified in part **a** react spontaneously in an electrochemical cell. Identify the anode and cathode.
- **c** Write an overall redox equation for the reaction.

4 Identify the conjugate redox pairs in each of the following redox reactions.

- **a** $H_2(g) + Cl_2(g) \rightarrow 2HCl(aq)$
- **b** $Cu(NO_3)_2(aq) + Fe(s) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$
- **c** $2AgNO_3(aq) + Pb(s) \rightarrow Pb(NO_3)_2(aq) + 2Ag(s)$

5 Identify the appropriate material for the electrodes in the following half-cells:

- **a** Zinc (Zn) half-cell where the electrode is the anode
- **b** Copper (Cu) half-cell where the electrode is the cathode
- **c** A solution of HNO_3 where the hydrogen ion undergoes reduction
- **d** Hydrogen telluride (H_2Te) is bubbled through water to create an acidified solution.

Design and discuss

6 Discuss why it is essential to separate a lithium/fluorine galvanic cell into two half-cells. Describe the consequences if the reactants are kept in the same beaker.

4.4

The electrochemical series

KEY IDEAS

In this topic, you will learn that:

- the electrochemical series is used to determine the relative strength of oxidising and reducing agents
- voltage produced by galvanic cells is calculated from standard electrode potentials.

Some chemicals will more readily gain or lose electrons than others. In Units 1 and 2, you learned how to use the periodic table and reactivity series of metals to predict redox reactions. However, this does not consider molecules or complex ions that contain more than one atom. A more accurate way of determining the strengths of oxidising or reducing agents is to use the **electrochemical series**.

electrochemical series
a table of oxidising agents and reducing agents written in reduction half-equations, where the strongest oxidising agent is at the top left of the table and the strongest reducing agent is at the bottom right

The electrochemical series

The electrochemical series is a table that shows all half-equations as reduction half-equations. It uses **equilibrium arrows** ($\rightleftharpoons$) to indicate that the half-equation can proceed forward or backward. When the half-equation goes forward, reduction occurs. When it goes backward, oxidation occurs. Oxidising agents are shown on the left side of the equilibrium arrow and reducing agents are shown on the right.

equilibrium arrows
reaction arrows that point in both directions to indicate that a reaction is reversible

The table orders chemical species based on their relative oxidising strength (strongest at the top left to weakest at the bottom left). They are all expressed as reduction half-equations.

F_2(g) is the strongest oxidising agent, because it will more readily accept electrons than any other molecule or ion and cause oxidation in another chemical species. The reduction half-equation is found at the top of the electrochemical series:

$$F_2(g) + 2e^- \rightarrow 2F^-(aq)$$

Li(s) is the strongest reducing agent, because it will more readily donate electrons than any other molecule or ion and cause reduction in another chemical species. The reduction half-equation for Li^+(aq) forming Li(s) is found at the bottom of the electrochemical series:

$$Li^+(aq) + e^- \rightarrow Li(s)$$

This half-reaction equation can be reversed to show the oxidation of Li(s):

$$Li(s) \rightarrow Li^+(aq) + e^-$$

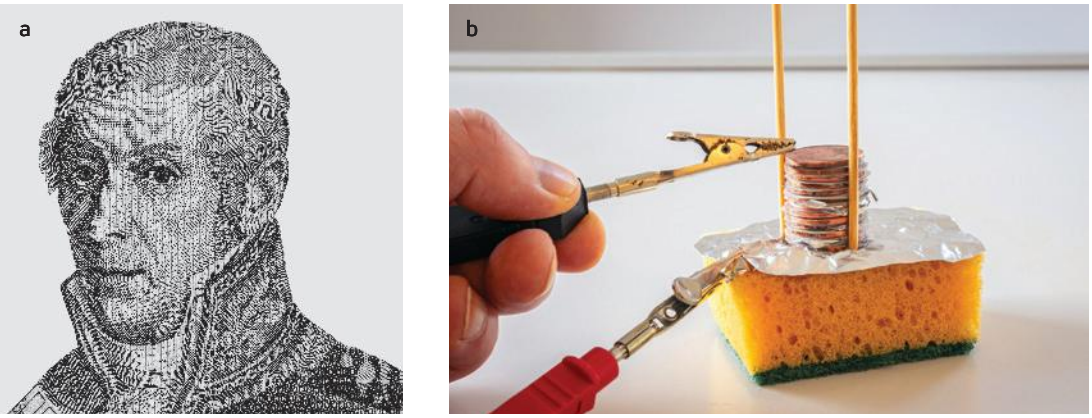


FIGURE 1 **a** Alessandro Volta was one of the first scientists to develop a ranking of metals based on their electrical effects. Volta also developed one of the first batteries in 1800. This stack of alternating copper and zinc disks, called a voltaic pile, is recreated in **b**.

4.4A
PRACTICAL: CONTROLLED EXPERIEMENT

How can we verify the electrochemical series using galvanic cells? Go to your **Student obook pro**.

4.4B
PRACTICAL: SIMULATION

Can you accurately predict cell voltage in a galvanic cell? Go to page **505**.

	Reaction	Standard electrode potential ($E^\ominus$) in volts at 25°C	
Strongest oxidising agents	$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87	Weakest reducing agents
↓	$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	+1.77	↑
	$Au^+(aq) + e^- \rightleftharpoons Au(s)$	+1.68	
	$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36	
	$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23	
	$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09	
	$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80	
	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77	
	$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	+0.68	
	$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54	
	$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40	
	$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34	
	$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15	
	$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(aq)$	+0.14	
	$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00	
	$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13	
	$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14	
	$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.25	
	$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28	
	$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	−0.40	
	$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.44	
	$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76	
	$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83	
	$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	−1.18	
	$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66	
	$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	−2.37	
Weakest oxidising agents	$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71	Strongest reducing agents
	$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	−2.87	
	$K^+(aq) + e^- \rightleftharpoons K(s)$	−2.93	
	$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.04	

FIGURE 2 The electrochemical series

Study tip

At the start of the writing time in your exam, make sure you annotate your electrochemical series with the strongest oxidising and reducing agents. This will be helpful for when you need to predict redox reactions.

Predicting redox reactions using the electrochemical series

To determine whether two chemicals will react spontaneously in a redox reaction, you can use the **RED SOX** method, which uses a diagram shaped as the letter ‘S’ (Figure 3).

1 Find the appropriate half-equations on the electrochemical series.
2 Identify the chemical species that are reacting.
3 Identify the strongest oxidising agent (top left), which will react spontaneously with the strongest reducing agent (bottom right). This is summarised in the ‘S’ configuration of Figure 3.
4 **Red**uction is on the top part of the **S** and **ox**idation is on the bottom, so the acronym is **RED SOX**. Rewrite the half-equations with the reactants and products on the correct sides of the reaction arrow.

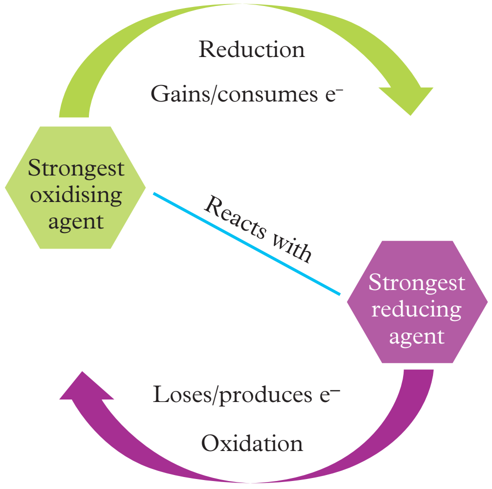


FIGURE 3 The ‘S’ configuration is used to determine whether chemical substances will react spontaneously.

RED SOX
an acronym to remember that reduction is the top half-equation on the electrochemical series, the S dictates the reactants and products, and the bottom half-equation is oxidation

Study tip

There are many methods you can use to identify reactants of a redox reaction; RED SOX and the ball-down-a-hill analogy are just two examples.

If the reactants do not fit in the ‘S’ configuration, they will not react. This occurs when the oxidising agent and reducing agent are not strong enough to react. Another analogy can be used to explain this: a ball will roll down a hill without any extra energy input. When the two reactants can be connected on the electrochemical series with a downhill line (left to right), then the reaction will be spontaneous. Compare this to RED SOX and see how this method is used to predict redox reactions in Worked example 4.4A.

4.4A WORKED EXAMPLE

PREDICTING REDOX REACTIONS USING THE ELECTROCHEMICAL SERIES

Predict whether a reaction will occur between:

a iron(II) nitrate ($Fe(NO_3)_2$) and magnesium metal (Mg)

b iron(II) nitrate ($Fe(NO_3)_2$) and lead(II) nitrate ($Pb(NO_3)_2$).

If a redox reaction does occur, write an overall redox equation for the reaction.

Solution

Think	Do
Step 1: Locate the relevant half-equations on the electrochemical series. It helps to write them in the order in which they appear in the electrochemical series. Iron(II) ions are involved in two half-equations.	**a** $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ $Mg^+(aq) + e^- \rightarrow Mg(s)$ **b** $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$
Step 2: Use RED SOX to help you determine whether a spontaneous redox reaction will occur. You can also use the rolling ball analogy.	**a** Fe^{2+} and Mg fit the ‘S’ configuration and can also be connected on the electrochemical series with a downhill line, so a spontaneous reaction will occur. **b** Fe^{2+} and Pb^{2+} do not fit the ‘S’ configuration and form an uphill line when connected on the electrochemical series, so a spontaneous reaction will not occur.
Step 3: Write the oxidation and reduction half-equations, then combine the two to make the overall redox equation. Go back to Worked example 4.2C if you need a refresher.	Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ Reduction: $Mg^+(aq) + e^- \rightarrow Mg(s)$ Overall equation: $Fe(s) + 2Mg^+(aq) \rightarrow Fe^{2+}(aq) + 2Mg(s)$

Study tip

After identifying the two redox half-equations on the electrochemical series, the top left and bottom right will react. If the reactants are not in this configuration, they will not react.

Measuring the strength of electrochemical cells

The position of all the half-reactions in the electrochemical series and their strength as oxidising and reducing agents are determined using the standard hydrogen half-cell ($H^+(aq)/H_2(g)$). This determines whether a half-reaction contains a stronger oxidising or reducing agent than the hydrogen cell.

- If reduction occurs in the hydrogen half-cell to form hydrogen gas (shown by bubbles in the solution), the hydrogen ion is the stronger oxidising agent and the other half-reaction sits below hydrogen on the electrochemical series.

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

You can also test the pH of the electrolyte solution. The consumption of H^+ in this reaction will decrease the H^+ ions in the solution. The cell will become less acidic, and the pH will increase.

- If oxidation occurs in the hydrogen half-cell, the hydrogen gas is the stronger reducing agent and the other half-reaction sits above hydrogen on the electrochemical series.

$$H_2(g) \rightarrow 2H^+(aq) + 2e^-$$

The production of H^+ will increase the concentration of H^+ ions in solution. The cell will become more acidic, and the pH will decrease.

Standard electrode potentials

In addition to observing the formation of hydrogen gas bubbles and measuring pH, you can also measure the voltage that is generated by the galvanic cell.

You might have noticed the E^0 values to the right of the electrochemical series in Figure 2. These are the **standard electrode potentials** – voltages generated by redox half-equations under **standard conditions** (where the temperature is 25°C, the pressure of a gas is 1 atm and the concentration of each solution is 1 M). Half-equations in the electrochemical series are ordered from highest to lowest E^0.

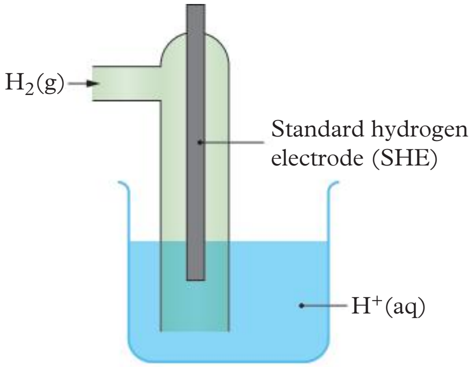


FIGURE 4 The standard hydrogen half-cell has an E^0 of 0.00 V.

The standard hydrogen half-cell is used as a reference cell to measure the voltages of all half-cells that are listed with an E^0 value on the electrochemical series. The half-cell in Figure 4 is given an arbitrary value of 0.00 V.

Calculating standard electrode potential

When a galvanic cell is constructed with a hydrogen half-cell and a second half-cell, the voltage generated by the cell is equal to the E^0 of the second half-cell. This is how the E^0 values on the electrochemical series have been measured.

In Figure 5, zinc generates a voltage of 0.76 V. Because this value is positive and we can see hydrogen gas bubbles being produced, we know that reduction is occurring in the hydrogen half-cell, so oxidation must be occurring at the zinc half-cell.

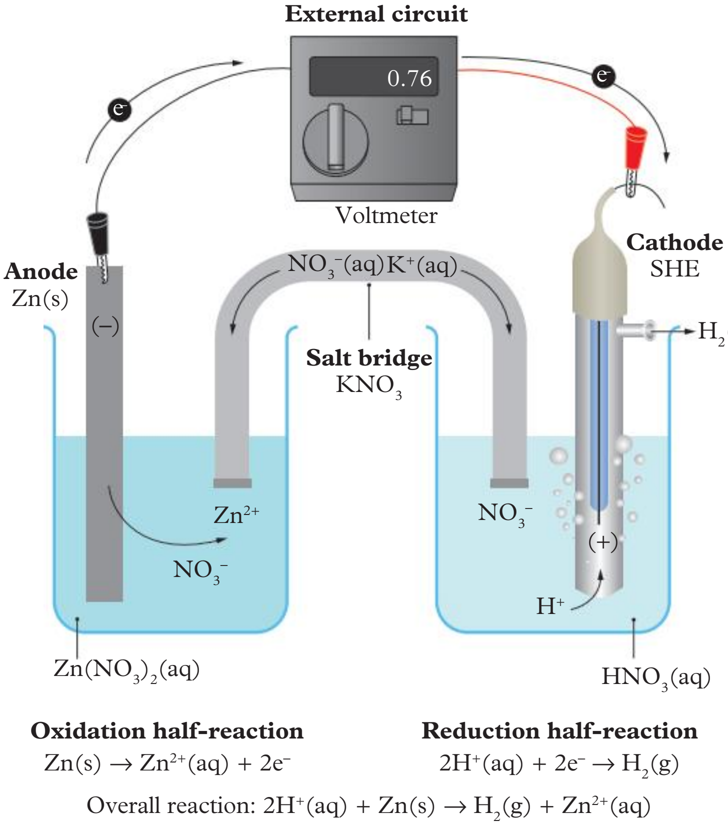


FIGURE 5 The zinc–hydrogen galvanic cell generates 0.76 V.

standard electrode potential
the electrical potential generated at the electrode under standard conditions; measured in volts (V)

standard conditions
the conditions that all E^0 values have been measured at; 25°C, 1 atm and 1 M

Study tip

Standard electrode potentials are measured at standard conditions of 25°C, 1 atm and 1 M concentration of solutions. The 0 in E^0 represents standard laboratory conditions.

Cell potential difference

When any two half-cells from the electrochemical series are connected in a galvanic cell, the voltage generated by the cell is the difference in the amount of electricity consumed at the cathode and generated at the anode – the **cell potential difference** (E^0_{cell}). This is also called the electromotive force (EMF, which indicates how much electricity is generated), the **standard cell potential** (if the cell is operating at standard conditions), or simply the voltage of the cell.

$$E^0_{\text{cell}} = E^0_{\text{reduction half-cell}} - E^0_{\text{oxidation half-cell}}$$

The E^0 is very important in the development of new battery technologies. It helps us choose materials that can generate a suitable maximum cell voltage when they react in a galvanic cell. You can see how to calculate E^0 in Worked example 4.4B.

cell potential difference
the difference in electrical potential between two electrodes, represented by E^0_{cell}; the voltage (in volts, V) that the galvanic cell can generate at standard conditions

standard cell potential
the same as cell potential difference

4.4B WORKED EXAMPLE

CALCULATING VOLTAGE PRODUCED BY A GALVANIC CELL

Calculate the maximum cell voltage produced when a Pb^{2+}/Pb half-cell is connected to an Al^{3+}/Al half-cell.

Solution

Think	Do
Step 1: Locate the relevant half-reactions on the electrochemical series. It helps to write them in the order in which they appear in the electrochemical series.	$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ $\quad E^0 = -0.13$ $Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ $\quad E^0 = -1.66$
Step 2: Use RED SOX to identify the oxidation and reaction half-equations.	$Pb^{2+}(aq)$ undergoes reduction to form $Pb(s)$. $Al(s)$ undergoes oxidation to form $Al^{3+}(aq)$.
Step 3: Substitute the E^0 values into the E^0_{cell} equation.	$E^0_{\text{cell}} = E^0_{\text{reduction half-cell}} - E^0_{\text{oxidation half-cell}}$ $= -0.13 - (-1.66)$ $= -0.13 + 1.66$ $= 1.53\,V$

Study tip

In galvanic cells, the reduction half-equation is higher on the electrochemical series than the oxidation half-equation. Therefore, E^0_{cell} = the top E^0 – the bottom E^0.

Limitations of the electrochemical series

When galvanic cells operate under non-standard conditions (i.e. a different temperature, pressure or concentration of solution) the E^0 values change. This can cause half-equations with close E^0 values to swap positions on the electrochemical series. This is a key limitation of the electrochemical series – we can only predict whether redox reactions will occur under standard conditions.

On top of this, the electrochemical series does not tell us about the rate of the reaction. You will learn about reaction rates and their importance in Unit 4 Area of Study 2.

FIGURE 6 A voltmeter or multimeter is often connected to a galvanic cell to measure the voltage produced in a redox reaction.

4.4 CHECK YOUR LEARNING

Describe and explain

1 Explain the purpose of the hydrogen half-cell in the electrochemical series.

2 Describe standard conditions. Explain what happens if a galvanic cell is not run at standard conditions, with reference to the electrochemical series.

Apply, analyse and compare

3 Calculate the E^0 of the following galvanic cells:

  a Cu^{2+}/Cu and Zn^{2+}/Zn

  b Mg^{2+}/Mg and Cu^{2+}/Cu

  c Fe^{2+}/Fe and Mg^{2+}/Mg

  d Cu^{2+}/Cu and Fe^{2+}/Fe

  e Mg^{2+}/Mg and Zn^{2+}/Zn

  f Zn^{2+}/Zn and Fe^{2+}/Fe

4 Consider the following half-equations:

$$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$$

$$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$$

$$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$$

  a Calculate the E^0 values for each combination of half-cells – there should be three combinations.

  b Draw fully labelled cell diagrams for the three galvanic cells from part **a**.

  c Determine which chemical species is the strongest oxidising agent and which is the strongest reducing agent.

  d Identify the combination of half-cells that produces the highest voltage.

Design and discuss

5 Discuss the impact of using a copper half-cell as the reference cell, rather than hydrogen. Describe the effect this would have on the voltage produced by the cell, the voltage calculated on the electrochemical series and the E^0 values of the electrochemical series.

6 Pieces of magnesium and zinc are often used to prevent solid iron from undergoing oxidation to form iron(II). Discuss why this occurs, making sure to reference oxidising agents, reducing agents and the electrochemical series.

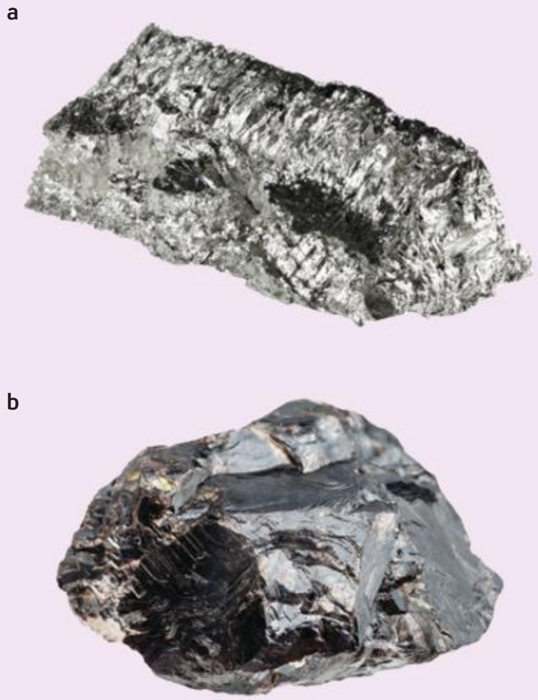


FIGURE 7 **a** Magnesium and **b** zinc

7 In a copper–silver galvanic cell, a student uses a sodium chloride salt bridge. After setting up the cell, she notices that it does not generate any electricity. Discuss why the cell produced no current. Evaluate her method and design an alternative that would work.

4.5 Features of fuel cells

KEY IDEAS

In this topic, you will learn that:

- fuel cells are galvanic cells that produce electrical energy from chemical energy by using a constant supply of reactants and catalyst-embedded porous electrodes
- fuel cells can be alkaline (basic) or acidic.

As you learned in Topic 4.3, primary galvanic cells such as those found in disposable batteries can only produce very small voltages. These are useful in many devices; unfortunately, it is not ideal to use disposable batteries to power something like a car.

fuel cell
a galvanic cell that produces electricity by using a constant supply of reactants (often hydrogen and oxygen) and inert electrodes that do not break down

Instead, **fuel cells** can be used. They are also a type of primary galvanic cell. They have a high power density (amount of power per unit of volume of reactant), are highly efficient (low amount of energy is lost in the conversion of chemical to electrical energy) and you can supply more reactant (fuel) to continue generating electrical energy. Although they are not rechargeable, they still save you from having to replace an expensive car battery every time the reactants are consumed.

Fuel cells

Fuel cells provide an alternative method for producing an electrical current. They are galvanic cells consisting of inert electrodes. They also require a constant supply of reactants.

combustion
a reaction occurring in the presence of oxygen to produce heat energy (exothermic)

In a fuel cell, the reactants are gaseous so there is no build-up of solid products that might block the flow of charge through the cell. A fuel cell can produce current for as long as it has a constant supply of reactants. Generally, the fuel is the reducing agent (which undergoes oxidation at the negative anode), and oxygen (O_2) is the oxidising agent (which undergoes reduction at the positive cathode). This is a **combustion** reaction.

Structure of a fuel cell

Reactants in a fuel cell must be contained within separate compartments to prevent them from reacting spontaneously with each other. Reactants enter the compartments through inlet channels. Excess materials exit through outlet channels and can be recycled back into the cell for use. In these compartments, the reactants come into contact with the electrodes (Figure 1).

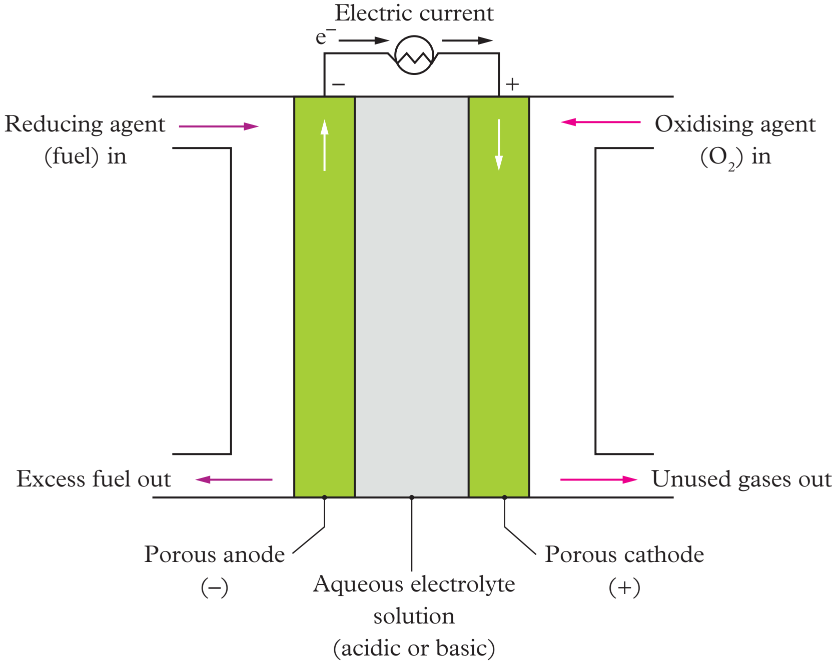


FIGURE 1 The general structure of a fuel cell

The electrodes are constructed from porous carbon. This increases their surface area and therefore provides more opportunities for the gaseous reactants to come into contact with the electrodes (Figure 2). They may also be embedded with a catalyst (black dots in Figure 2), which increases the rate of the reaction and therefore the efficiency of the cell. You will learn about how catalysts work in Chapter 5. The electrodes are connected by a wire that allows electron movement.

The electrolyte is in contact with both electrodes. The nature of the electrolyte (acidic or alkaline) has a significant effect on the half-reactions occurring within the fuel cell. In alkaline (or basic) fuel cells, a basic substance, commonly a hydroxide (KOH), is used as an electrolyte. Acidic fuel cells use an acidic electrolyte, such as the strong sulfuric acid (H_2SO_4) or hydrochloric acid (HCl).

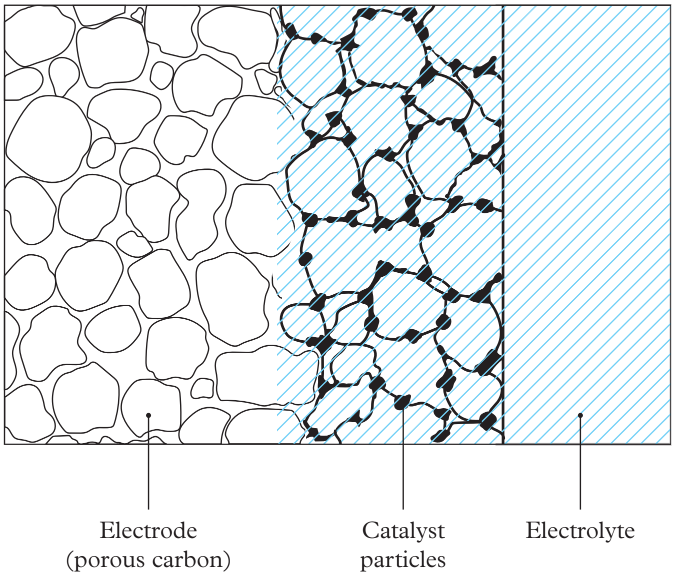


FIGURE 2 Electrodes are made of porous carbon and often have catalyst particles embedded in them.

Hydrogen–oxygen fuel cells

To understand how electrolyte solutions affect the oxidation and reduction equations in a fuel cell, let's use the hydrogen–oxygen fuel cell as an example. This is one of the most common types of fuel cells.

Looking at the electrochemical series, there are four possible half-equations that use hydrogen (H_2) and oxygen (O_2):

1 $O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l) \quad E^0 = +1.23$

2 $O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq) \quad E^0 = +0.40$

3 $2H^+(aq) + 2e^- \rightleftharpoons H_2(g) \quad E^0 = 0.00$

4 $2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq) \quad E^0 = -0.83$

Which of these half-equations occur will depend on the nature of the electrolyte.

In acidic conditions, half-equations involving $H^+(aq)$ are included (equations 1 and 3). In alkaline conditions, half-equations involving $OH^-(aq)$ are included (equations 2 and 4). As fuel cells involve spontaneous redox reactions, the strongest oxidising agent is oxygen gas ($O_2(g)$) and the strongest reducing agent is hydrogen gas ($H_2(g)$). Remember RED SOX!

Regardless of whether an acidic or basic electrolyte is used, both generate a standard cell potential of 1.23 V (Table 1).

TABLE 1 Half-equations, overall balanced equations and voltage generated by acidic and alkaline fuel cells at standard conditions

	Acidic fuel cell	**Alkaline fuel cell**
Reduction (cathode)	$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
Oxidation (anode)	$H_2(g) \rightarrow 2H^+(aq) + 2e^-$	$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$
Overall	$O_2(g) + 2H_2(g) \rightarrow 2H_2O(l)$	$O_2(g) + 2H_2(g) \rightarrow 2H_2O(l)$
Voltage at standard conditions	E^0 = reduction − oxidation = +0.40 − (−0.83) = 1.23 V	E^0 = reduction − oxidation = +1.23 − (0.00) = 1.23 V

Study tip

The RED SOX method can be used since fuel cells involve spontaneous redox reactions. If a fuel cell contains a reactant that is not on the electrochemical series, identify the reactant and product and then use the KOHES method to complete the half-equation.

Study tip

Always start by identifying the reactants and products. Then, determine whether they are undergoing oxidation or reduction and write the half-equations. Balance the number of electrons, then combine to form the overall equation.

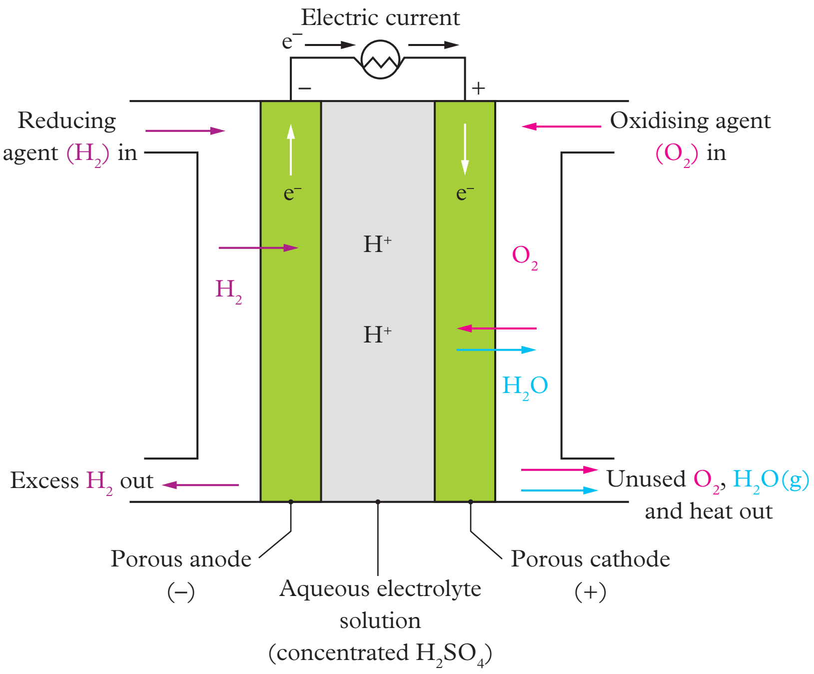


FIGURE 3 An acidic fuel cell with an electrolyte that supplies H^+ ions

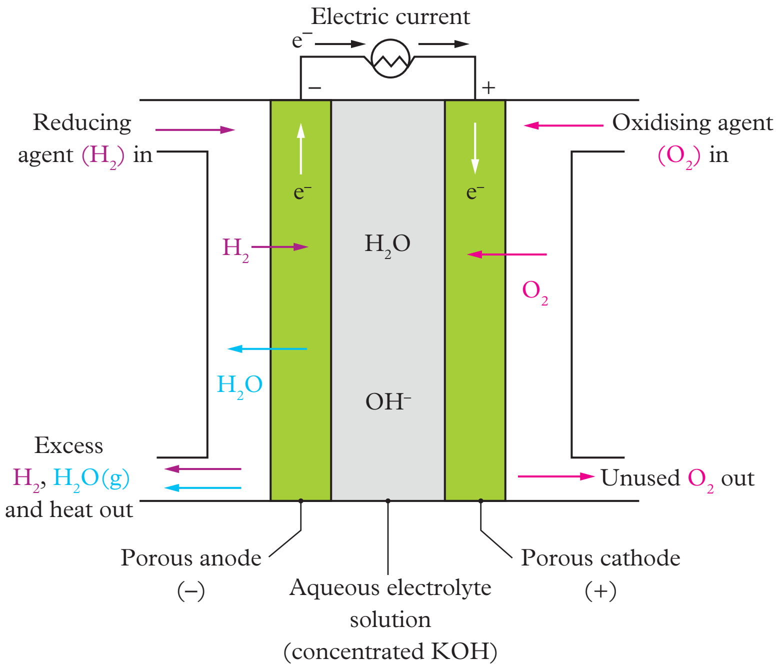


FIGURE 4 An alkaline fuel cell with an electrolyte that supplies OH^- ions

Acidic fuel cells

In a hydrogen–oxygen fuel cell, hydrogen enters the acidic fuel cell at the negative anode and is oxidised to produce hydrogen ions and electrons. The electrons move to the positive cathode via the external circuit (wires) while the hydrogen ions migrate from the anode to the cathode via the internal circuit (the electrolyte).

At the cathode, the reduction of oxygen occurs. Hydrogen ions (from the electrolyte) and oxygen molecules, which are in contact with the electrode, gain electrons to produce water molecules. Excess reactants are captured and recycled back into the fuel cell while the products of water and heat are released into the atmosphere.

Alkaline fuel cells

Hydrogen enters the alkaline fuel cell at the negative anode. At the anode, hydrogen undergoes oxidation to produce water, hydroxide ions and electrons. The electrons move to the positive cathode via the external circuit (wires) while the water migrates from the anode to the cathode via the internal circuit (the electrolyte).

At the cathode, the reduction of oxygen occurs. Water (from the electrolyte) and oxygen molecules gain electrons to produce hydroxide ions in the electrolyte. These ions move towards the anode through the electrolyte. Excess reactants are captured and recycled back into the fuel cell while the products of water and heat are released into the atmosphere.

Hydrocarbon–oxygen fuel cells

Fuel cells do not always use hydrogen and oxygen gases as reactants. Other hydrocarbon fuels can be used, such as methane. Methane reacts with oxygen in the following combustion reaction:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

To create half-equations from this reaction, methane is paired with carbon dioxide as the carbon containing molecules must remain in the same half-equation. Therefore, oxygen is paired with water in the second half-equation.

The oxygen molecule, $O_2(g)$, has an oxidation number of 0. In water, the oxidation number of oxygen is –2. As the oxidation number decreases, $O_2(g)$ undergoes reduction. This is true for all combustion reactions involving hydrocarbons. Oxygen is reduced and the hydrocarbon is oxidised.

The oxidation half-equation for the combustion of methane is therefore:

$$CH_4(g) + 2H_2O(l) \rightarrow CO_2(g) + 8H^+(aq) + 8e^-$$

The reduction half-equation is:

$$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$$

FIGURE 5 The combustion of hydrocarbon fuels powers many vehicles

Study tip

Again, fuel cells (a type of galvanic cell) can be connected to a *load* that uses the generated electrical energy.

4.5 REAL-WORLD CHEMISTRY

Launching a space shuttle

Fuel cells are commonly used in space shuttles. Each space shuttle contains three fuel cell power plants, which are supplied with hydrogen and oxygen gas. The water produced is stored and used for other purposes. Any heat produced by the fuel cell is transferred to freon coolant loops. The electricity generated is directed toward the shuttle's power systems.

Tanks external to the shuttle are filled with liquid hydrogen and liquid oxygen. They help to launch the space shuttle into orbit. When hydrogen and oxygen react spontaneously, they produce large quantities of heat energy in an exothermic reaction. The direct reaction also generates light and sound energy. The hydrogen and oxygen are then separated in a galvanic (fuel) cell, where they produce the far more beneficial electrical energy.

FIGURE 6 The space shuttle attached to external fuel tanks

Approximately 8.5 minutes into launch, the fuel is completely consumed. The external tanks become useless and are now, essentially, dead weight. The tanks detach from the space shuttle and fall back to Earth.

Apply your understanding

1 Write the oxidation and reduction half-equations for this alkaline fuel cell.

2 Write the overall balanced chemical equation for this cell.

3 The external tanks used to propel the shuttle into space use a redox reaction but do not produce electrical energy. Explain why.

4.5 CHECK YOUR LEARNING

Describe and explain

1 Describe the difference between a primary cell and a fuel cell. Give examples of both.

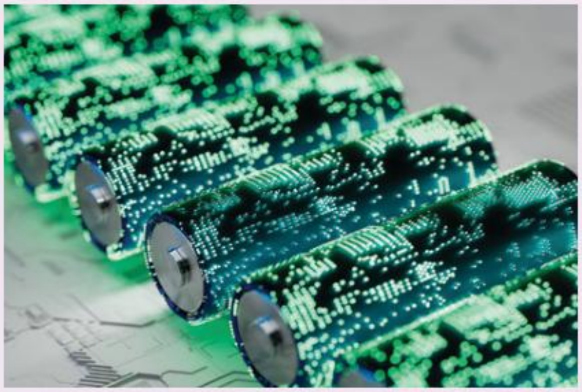


FIGURE 7 Primary and fuel cells both generate electrical energy.

2 Explain why porous electrodes are used in fuel cells.

Apply, analyse and compare

3 Construct a diagram of an acidic ethanol–oxygen fuel cell which produces carbon dioxide and water as products. In your answer, you must:

- **a** construct two half-equations and identify them as oxidation and reduction
- **b** indicate the reactants, products and electrolyte
- **c** label the anode and cathode
- **d** use arrows to show the direction of movement of the electrons and the ions in the electrolyte
- **e** combine the half-equations into a balanced overall equation.

4 Construct a diagram of an alkaline methanol–oxygen fuel cell which produces carbon dioxide and water as products. In your drawing, you must:

- **a** construct two half-equations and identify them as oxidation and reduction
- **b** indicate the reactants, products and electrolyte
- **c** label the anode and cathode
- **d** use arrows to show the direction of movement of the electrons and the ions in the electrolyte
- **e** combine the half-equations into a balanced overall equation.

5 Compare fuel cells and galvanic cells in terms of:

- their set-up
- how they work
- what they are used for
- specific design features
- energy transformation
- energy efficiency
- applications.

FIGURE 8 The car battery is an example of a fuel cell.

4.6 Faraday's Laws

KEY IDEAS

In this topic, you will learn that:

- Faraday's Laws can be used to predict the quantity of reactant and product, and the amount of current or time required to consume a certain quantity of reactant or produce a certain quantity of product
- Faraday's first law states that the charge of a cell is proportional to the mass plated at the cathode
- Faraday's second law states that in order to produce 1 mole of a metal, 1, 2, 3 or another whole number of moles of electrons must be consumed.

Techniques for quantitative analysis of electrochemical cells were first developed by Michael Faraday in 1833. Faraday developed his laws using a process called electrolysis, whereby electrical energy is applied to an electrochemical cell to reverse a spontaneous chemical reaction. This converts the products of a spontaneous reaction back into reactants so that they can be used again to generate electrical energy. You will learn about electrolysis in Chapter 7.

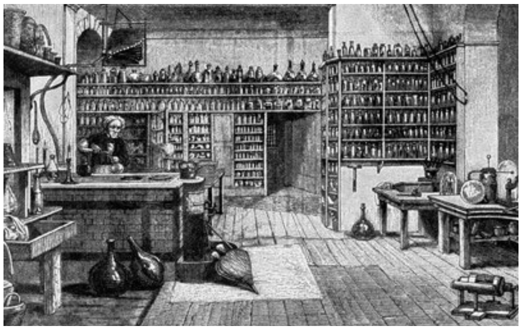


FIGURE 1 Faraday's laboratory

In this topic, we will use the laws defined by Faraday to make predictions about the amounts of reactants used and products formed in redox reactions.

Faraday's Laws

In an **electroplating cell**, a metal will be deposited onto the cathode of the cell. Faraday's Laws describe the relationship between the mass of the metal that is deposited and the amount of charge that is applied to the cell.

electroplating cell
an electrolytic cell that converts electrical energy to chemical energy in order to coat an object in a layer of metal

Faraday's first law

Current is a flow of charge created by the movement of electrons. To measure the amount of charge flowing through an electrochemical cell, an ammeter must be attached in the external circuit. An ammeter measures the **current** of the cell in amps (A). A timer is also required to determine the length of time (in seconds) that the charge is applied for.

current
the flow of electrons generated by a power source; measured in amps (A) and represented by *I*

The amount of **electric charge** which passes through a cell can be calculated using:

$$Q = I \times t$$

where Q is the electric charge in coulombs (C), I is the current in amps (A) and t is the time in seconds (s).

electric charge
the rate of flow of electrons generated by a power source; measured in coulombs (C) and represented by *Q*

Faraday found that increasing the charge will increase the mass of metal which plates onto the cathode. Essentially, the higher the current and the longer it runs, the larger the mass of metal or volume of gas produced at the cathode. You can see the relationship between the mass of different metals and the charge applied to the cell in Figure 2 on the next page.

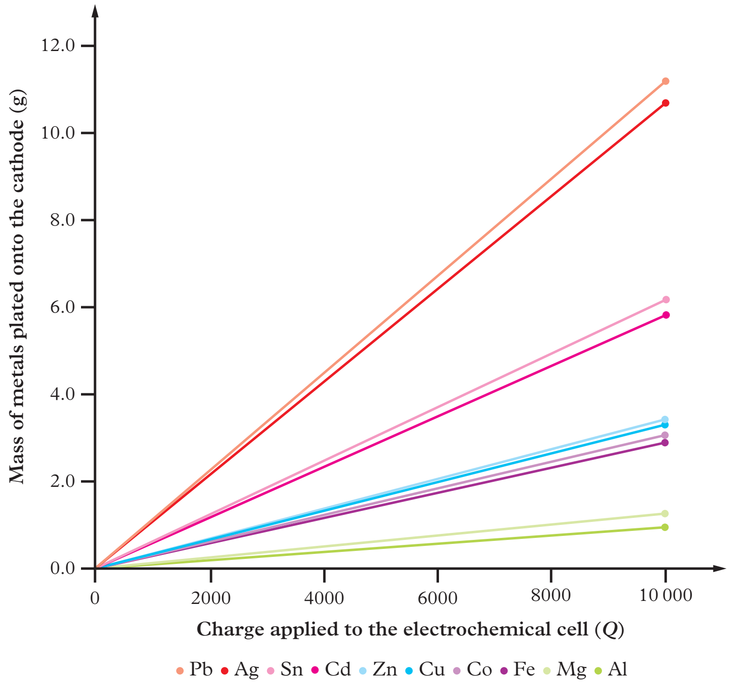


FIGURE 2 The mass of various metals plated at the cathode when charge is applied to an electrolytic cell.

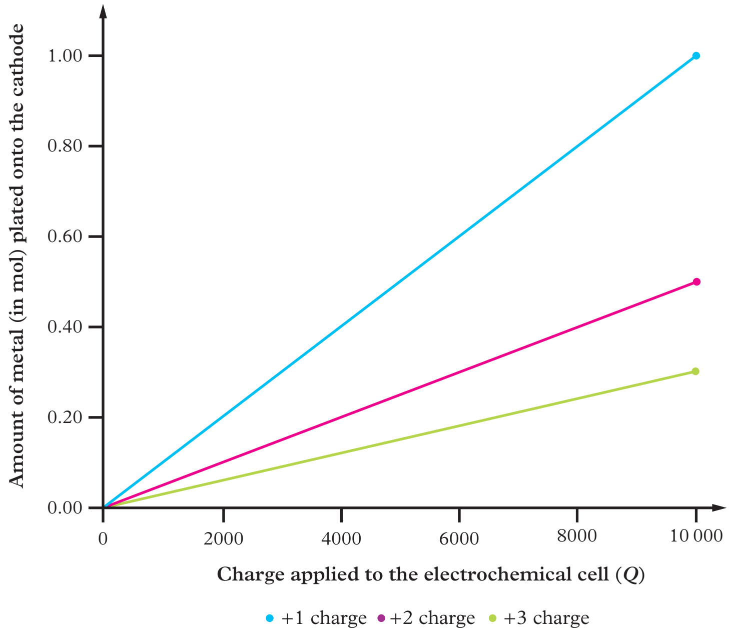


FIGURE 3 The amount, in mol, of various metals plated at the cathode of an electrolysis cell when charge is applied to an electrolytic cell

Faraday's first law
the mass of a metal plated onto a cathode is proportional to the charge applied to an electroplating cell

From these observations, Faraday determined that the mass of the metal plated onto the cathode is proportional to the charge applied to the cell. This is **Faraday's first law**:

$$Q \propto m$$

where Q is the electric current and m is the mass of the metal deposited onto the cathode.

Faraday's second law

Faraday then converted the mass of each metal into moles by dividing each mass by the metal's molar mass. He plotted the amount, in mol, of each metal plated at the cathode against the charge applied to the cell. Figure 3 shows that similar trends were observed for metals ions with the same charge (or electrovalency).

All metal ions with a +1 charge had the same linear trend (e.g. Na^+, K^+, Ag^+). The same was true of metal ions with +2 (e.g. Ca^{2+}, Cu^{2+}) and +3 charges (e.g. Fe^{3+}).

Faraday also found that:

- all +1 metals required 96 500 coulombs of charge to plate 1 mol of the metal onto the cathode.
- all +2 metals required 2 × 96 500 coulombs of charge to plate 1 mol of metal onto the cathode. Therefore 0.5 mol of electrons were plated using 96 500 C.
- all +3 metals required 3 × 96 500 coulombs of charge to plate 1 mol of metal onto the cathode. Therefore 0.33 mol of electrons were plated using 96 500 C.

Faraday concluded that the charge on one mole of electrons must be 96 500 coulombs. This is known as the **Faraday constant**:

$$F = 96\,500\,\text{C mol}^{-1}.$$

This means that 1, 2, 3 or another whole number of moles of electrons must be consumed to produce 1 mol of a metal. This is **Faraday's second law**:

$$Q = n(e^-) \times F \text{ or } Q = n(e^-) \times 96\,500$$

where Q is the electric current, $n(e^-)$ is the amount, in mol, of electrons and F is the Faraday constant.

Charge on one electron

The charge of one mole of electrons is 96 500 C. If you consider that one mole of electrons is 6.02×10^{23} electrons, then the charge on one electron can be calculated as:

$$\frac{96500}{6.02 \times 10^{23}} = 1.603 \times 10^{-19}\,\text{C}$$

As electrons are negatively charged, this number is typically represented as -1.60×10^{-19} C.

Faraday constant
1 F = 96 500 C mol^{-1}; 96 500 coulombs are required to generate one mole of a metal or gas at the cathode of an electroplating cell

Faraday's second law
to produce 1 mole of a metal, 1, 2, 3 or another whole number of moles of electrons must be consumed

The application of Faraday's Laws

Although Faraday's Laws were developed using electroplating cells, which consume electrical energy, we can also apply them to galvanic and fuel cells, which generate electrical energy.

These laws can be used in calculations to determine the life of a battery, the length of time that it takes for a cell to consume all of its reactants and the amount of reactant required to ensure that a cell lasts for a desired period of time.

Calculating the quantity of reactants and products

The amount of reactant in a battery or cell will decrease over time as the amount of product increases. This can be calculated as a decrease in the mass of reactants at a reactive electrode or a decrease in the concentration of aqueous reactants. Alternatively, it can be the increase in the mass of an electrode when the electrolyte plates onto it, or an increase in the concentration of aqueous products.

Worked example 4.6A shows how to calculate the mass of reactants used in a galvanic cell. You can use the same steps to determine the mass of products formed.

Study tip

Remember that you can calculate the amount, in mol, of a gas at SLC from volume using the molar volume equation:

$$n_{\text{SLC}} = \frac{V}{24.8}$$

4.6A WORKED EXAMPLE

CALCULATING THE QUANTITY OF REACTANTS USED IN A GALVANIC CELL

A copper–silver galvanic cell consisting of 1 M concentrations of Cu/Cu^{2+} and Ag/Ag^{+} is connected to a small 40 W light bulb for 75.0 minutes. The average current drawn by the bulb is 0.36 A. Calculate the mass of copper that is oxidised during this time.

Solution

Think	Do
Step 1: Calculate the charge produced by the galvanic cell. Remember that time needs to be in seconds.	$Q = I \times t$ $= 0.36 \times (75.0 \times 60)$ $= 1620$ C
Step 2: Calculate the amount, in mol, of electrons consumed.	$n(e^-) = \frac{Q}{F}$ $= \frac{1620}{96\,500}$ $= 0.0168$ mol
Step 3: Write a balanced half-equation for the half-reaction.	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
Step 4: Use a ratio statement to determine the amount, in mol, of reactant used.	$n(Cu) = \frac{1}{2} \times n(e^-)$ $= 0.00839$ mol
Step 5: Calculate the mass of reactant used.	$m(Cu) = n \times M$ $= 0.00839 \times 63.5$ $= 0.52$ g (2 sig fig)

Study tip

If a half-equation is not given in the question, it should be included in your answer with the ratio statement. If it is given in the question, there is no need to rewrite it in your answer – just write the ratio statement.

Calculating quantities in batteries

To calculate the time required to consume all the reactants in a galvanic cell (the battery life of a cell) or the time required to use a particular quantity of reactant, you need to know how much reactant is present, and the current that is produced. Worked example 4.6B walks you through how to do this.

4.6B WORKED EXAMPLE

CALCULATING THE LIFE OF AN AA ALKALINE BATTERY

A typical AA alkaline battery can produce up to 2.8 A. The battery contains 6.5 g of MnO_2. Calculate the time it takes, in minutes, for the battery to die if it runs constantly at maximum capacity.

$$Zn(s) + 2MnO_2(s) + H_2O(l) \rightarrow Zn(OH)_2(s) + Mn_2O_3(s)$$

Solution

Think	Do
Step 1: Calculate the amount, in mol, of MnO_2.	$n(MnO_2) = \frac{m}{M}$ $= \frac{6.5}{86.9}$ $= 0.075\,mol$
Step 2: Write a balanced half-equation for the reduction of MnO_2.	$2MnO_2(s) + H_2O(l) + 2e^- \rightarrow Mn_2O_3(s) + 2OH^-(aq)$
Step 3: Use a ratio statement to determine the amount, in mol, of electrons consumed.	$n(e^-) = n(MnO_2)$ $= 0.075\,mol$
Step 4: Calculate the charge.	$Q = n(e^-) \times F$ $= 0.075 \times 96\,500$ $= 7218.1\,C$
Step 5: Calculate the time. Remember that time is in seconds. Here, you will need to convert to minutes.	$t = \frac{Q}{I}$ $= \frac{7218.1}{2.8}$ $= 2577.88\,sec$ $2577.88 \div 60 = 43\,min$ (2 sig fig)

4.6 SKILL DRILL

Designing an experiment to verify Faraday's second law

Key science skill: Plan and conduct investigations

A student wants to verify Faraday's second law of electrolysis. She compares each of the following half-cells to the standard hydrogen half-cell (H_2/H^+):

Ag/Ag^+ Au/Au^+ Cu/Cu^{2+} Mg/Mg^{2+}

She applies 6 V to each electrochemical cell for 5.0 minutes and generates a current of 0.25 A in each cell.

Practise your skills

1 Identify the independent and dependent variables.
2 Write a hypothesis for the experiment.
3 Identify the aim of the experiment.

Need help planning and conducting experiments? See Topic 1.4 (page 14).

4.6 CHECK YOUR LEARNING

Describe and explain

1 Explain what 1 faraday represents.

2 Explain what electric charge is and how it can be calculated.

3 Explain how the number of moles of a metal or ion can be determined using the number of moles of electrons.

Apply, analyse and compare

4 A button cell is used to power a 60 W light bulb requiring a current of 0.54 A. If the cell originally contains 0.50 g of zinc metal, calculate how long it will take to convert 10% of the metal into zinc hydroxide.

$$Zn(s) + Ag_2O(s) + H_2O(l) \rightarrow 2Ag(s) + Zn(OH)_2(s)$$

5 A cordless vacuum cleaner battery can provide 18 V for 7 hours. The vacuum cleaner uses a lithium–ion battery with a CoO_2 mass of approximately 10.0 g.

$$CoO_2(s) + Li^+(aq) + e^- \rightarrow LiCoO_2(l)$$

$$LiC_6(s) \rightarrow C_6(s) + Li^+(aq) + e^-$$

If the mass of CoO_2 decreases by 1.5 g after every use, and typical use lasts for 35 minutes, calculate the current generated by the battery.

6 A methanol fuel cell uses an acidic electrolyte. 1.50 L of methanol (density = 0.792 g mL^{-1}) is combusted into carbon dioxide when pumped through the cell over the course of 4.500 hours. The half-equation for this reaction has an E^0 value of −0.39 V. The cell can produce a current of 85 A.

- **a** Calculate the volume of methanol that is consumed in 4.500 hours.
- **b** Calculate the percentage of methanol remaining for future use.

Design and discuss

7 A laptop uses a rechargeable lithium–ion battery (see Question 5). Initially, the battery charges to 100% capacity, but 12 months later it can only charge to 80% capacity.

- **a** Discuss what is meant by the term *capacity* in the context of electrochemistry.
- **b** If the current produced by the battery at 100% capacity is 2.8 A, determine the current produced at 80%.
- **c** If a battery typically takes 7 hours to discharge fully, calculate the mass loss of LiC_6 when the laptop is being used at full power and at 100% capacity.

8 A scientist is attempting to construct a lithium–fluorine battery. The battery must be capable of a voltage of 5.5 V and must maintain a current of 30 A for a minimum of 8.0 hours.

- **a** Write the two half-equations and a balanced overall equation for the battery.
- **b** Calculate the standard cell potential and compare it to 5.5 V. Determine whether an adequate voltage is produced.
- **c** Calculate the mass, in g, of lithium metal required for this cell.
- **d** If the cell operates under SLC, calculate the volume of fluorine gas required.
- **e** Outline the safety risks of constructing this cell. Explain what redox reactions you would suggest instead of these.

4.7 Meeting society's energy needs

KEY IDEAS

In this topic, you will learn that:

- fuel cells can be constructed from renewable feedstocks
- fuel cells can be designed to be more efficient.

With global populations increasing, there is more demand than ever for energy to meet the needs of Earth's inhabitants. Combined with the climate crisis we are facing, it has made us rethink where we source energy.

The current and future options for supplying energy are the core focus of Unit 3 Area of Study 1. In this topic, we will explore how the green chemistry principles have guided the design of fuel cells, and how we hope this will address some of the challenges we face.

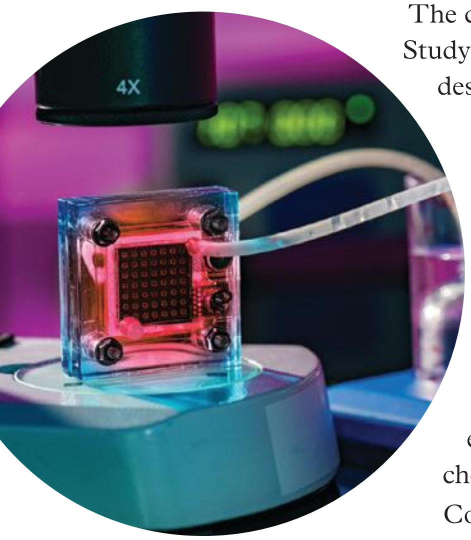


FIGURE 1 A hydrogen fuel cell

Innovation in the design of fuel cells

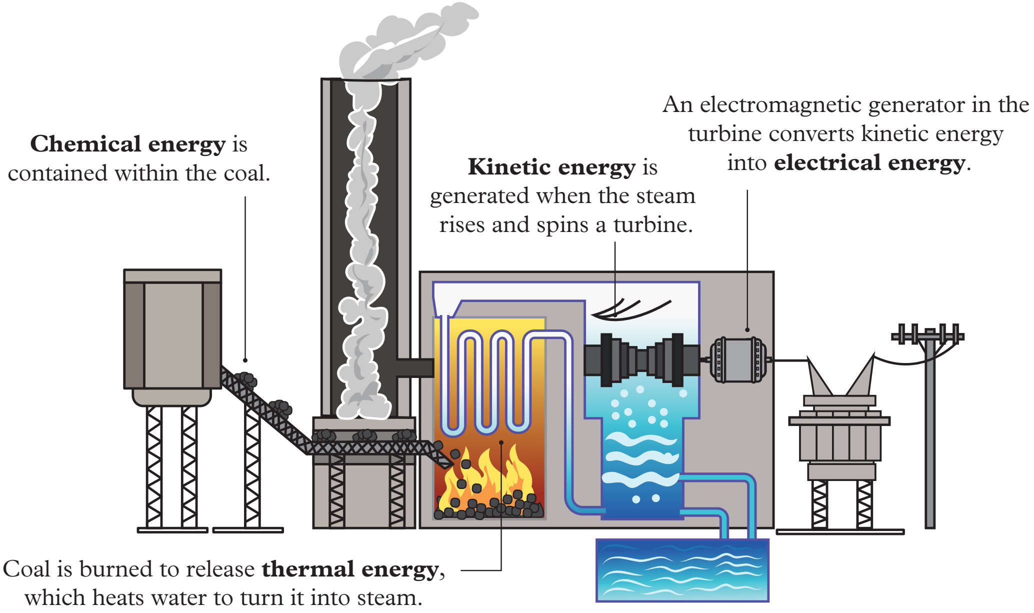
The design of fuel cells has incorporated many green chemistry principles. They are energy-efficient due to their porous and catalysed electrodes, and can be designed to use fuels from sustainable sources.

Design for energy efficiency

Fuel cells also convert chemical energy directly into electrical energy, with some energy wasted as heat. They are 60% efficient, which means that about 40% of the chemical energy is lost as heat.

Compare this to a combustion-based power plant, such as a coal-fired power station. The process involved in generating electrical energy from coal is shown in Figure 2. A coal power station runs at approximately 33–35% efficiency, due to the number of energy transformations that occur in the power station. More energy is lost as waste, and this decreases the efficiency of the process.

FIGURE 2 The process of turning chemical energy from coal into electricity in a coal-fired power station



Why are hydrogen fuel cells not widely used? Go to your **Student obook pro**.

Design and applications of fuel cells also incorporate the green chemistry principle of design for energy efficiency, because their chemical reactions can occur at room temperature and pressure. The use of porous electrodes increases the reaction rate. Addition of catalysts also does this, without the need to increase temperatures or pressures. You will learn more about reaction rates in Chapter 5.

Renewable feedstocks

As a society, we need to work towards reducing (or even eliminating) our use of non-renewable fuels. The burning of fossil fuels at coal-fired power stations releases **by-products** such as carbon dioxide and water vapour. These gases contribute to the **greenhouse effect** and **climate change**.

To address these environmental challenges, we need to design alternative fuel cells to meet society's energy needs. One approach is to use renewable feedstocks. The green chemistry principle refers to the ability to use starting materials that are renewable.

In fuel cells, renewable feedstocks are often sourced from plants and biomass as an alternative to fossil fuels such as crude oil or natural gas.

FIGURE 3 Biomass-derived fuels are renewable.

by-product
a secondary product formed in a chemical reaction

greenhouse effect
a phenomenon whereby greenhouse gases (e.g. methane, carbon dioxide, water vapour) trap heat in the atmosphere and cause the Earth to warm

climate change
the gradual increase in the temperature of the Earth's surface, oceans and atmosphere; generally caused by the greenhouse gas effect

In Topic 4.5, you learnt that the most common fuels used in fuel cells are hydrogen gas, oxygen gas and various hydrocarbons, as outlined in your data book. Table 1 below summarises the renewable fuel alternatives that can be used in fuel cells instead of the traditional non-renewable fuels.

TABLE 1 Sources of renewable and non-renewable fuels

Renewable fuels		Non-renewable fuels	
Source	**Fuels**	**Source**	**Fuels**
Biomass	Hydrogen, methane	Natural gas	Methane
Plants	Bioethanol	Crude oil	Hydrocarbons

Have a go at designing a green fuel cell in Challenge 4.7.

4.7 Challenge
Find me in your obook pro

4.7 CHECK YOUR LEARNING

Describe and explain

1 Describe the current challenges faced by society in meeting energy needs.

2 Explain why generating electrical energy in a fuel cell is more efficient than in a coal-fired power plant.

Design and discuss

3 Discuss how green chemistry principles can be used to guide the fuel choice for a fuel cell.

4 A fuel cell captures excess reactants and recycles them for use in the cell. Discuss how this is beneficial, with reference to the relevant sustainability perspectives (green chemistry principles, sustainable development goals, and the shift from a linear economy to a circular economy).

5 Chapter 2 explores the generation of fuels from sustainable and unsustainable resources. Design an experiment that generates fuels which can be used in a fuel cell. Your answer could explain the generation, capture and safety involved with hydrogen, methanol, ethanol or any other fuel from your VCE Chemistry data book.

6 Identify any other green chemistry principles (aside from design for energy efficiency and use of renewable feedstocks) that relate to using fuel cells rather than primary cells. Justify your response.

CHAPTER 4 Review

Chapter summary

4.1

- A redox reaction involves reduction (gain of electrons) and oxidation (loss of electrons) reactions, which occur simultaneously.
- Oxidation numbers show the transfer of electrons from a reducing agent to its conjugate oxidising agent, or from an oxidising agent to its conjugate reducing agent.

4.2

- Balanced oxidation and reduction half-equations can be written for redox reactions. Half-equations can be combined to form overall redox equations.
- Balanced complex oxidation and reduction half-equations and complex overall redox equations are written for reactions occurring in acidic and basic conditions, using the KOHES method.

4.3

- Redox reactions are used to generate electricity in a galvanic cell, where an external circuit connects two half-cells to facilitate electron flow.
- Reduction occurs at the positive electrode, called the cathode. Oxidation occurs at the negative electrode, called the anode.
- The internal circuit consists of a salt bridge, which is made of an unreactive salt solution. Negative ions move towards the negative electrode to balance the positive ions that move into the solution when the metal undergoes oxidation. Positive ions move towards the positive electrode to replace the positive ions that plate as a metal onto the cathode when they undergo reduction.
- A non-rechargeable battery is called a primary cell. There are many different types of primary cells – for example, alkaline batteries and button cells.

4.4

- The electrochemical series is used to determine the strength of oxidising and reducing agents.
- The voltage generated by a galvanic cell can be determined by using the electrochemical series. The oxidation E^0 is subtracted from the reduction E^0 to calculate the cell potential difference (also called the EMF or voltage).

4.5

- Fuel cells are galvanic cells that produce electrical energy from chemical energy by using a constant supply of reactants. They used catalyst-embedded porous electrodes to increase the rate of the reaction.
- Fuel cells use alkaline (basic) or acidic electrolytes. The nature of the electrolyte must be considered in half-equations.

4.6

- Faraday's Laws can be used to predict the quantity of reactant and product or the amount of current or time required to consume a certain quantity of reactant or produce a certain quantity of product.
- Faraday's first law states that the charge of a cell is proportional to the mass plated at the cathode.
- Faraday's second law states that in order to produce 1 mole of a metal, 1, 2, 3 or another whole number of moles of electrons must be consumed.

4.7

- Fuel cells are designed to incorporate green chemistry principles – they are designed for energy efficiency and can be constructed from renewable feedstocks.

Key formulas

Electric charge	$Q = I \times t$
Amount, in mol, of electrons	$n(e^-) = \frac{Q}{F}$
Standard cell potential (V)	E^0 = reduction E^0 − oxidation E^0
Faraday's first law	$Q \propto m$
Faraday's second law	$Q = n(e^-) \times F$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Mostly	Not really	Revision link
• explain that redox reactions involve simultaneous oxidation and reduction processes	☐	☐	☐	Go back to Topic 4.1
• use oxidation numbers to identify the reducing agent, oxidising agent and conjugate redox pairs	☐	☐	☐	Go back to Topic 4.1
• write balanced half-equations (including states) for oxidation and reduction reactions, and the overall redox cell reaction in both acidic and basic conditions	☐	☐	☐	Go back to Topic 4.2
• describe the common design features and general operating principles of non-rechargeable (primary) galvanic cells converting chemical energy into electrical energy, including electrode polarities and the role of the electrodes (inert and reactive) and electrolyte solutions	☐	☐	☐	Go back to Topic 4.3
• describe the use and limitations of the electrochemical series in designing galvanic cells and as a tool for predicting the products of redox reactions, for deducing overall equations from redox half-equations and for determining maximum cell voltage under standard conditions	☐	☐	☐	Go back to Topic 4.4
• describe the common design features and general operating principles of fuel cells, including the use of porous electrodes for gaseous reactants to increase cell efficiency	☐	☐	☐	Go back to Topic 4.5
• apply Faraday's Laws and stoichiometry to determine the quantity of galvanic or fuel cell reactant and product, and the current or time required to either use a particular quantity of reactant or produce a particular quantity of product	☐	☐	☐	Go back to Topic 4.6
• describe contemporary responses to challenges and the role of innovation in the design of fuel cells to meet society's energy needs, with reference to the green chemistry principles: design for energy efficiency and use of renewable feedstocks	☐	☐	☐	Go back to Topic 4.7

Revision questions

Multiple choice

The following equation relates to Questions 1 and 2.

$2Au^{3+}(l) + 6Br^{-}(l) \rightarrow 2Au(s) + 3Br_2(l)$

1 In this reaction:
- **A** Au is oxidised; Br_2 is reduced.
- **B** Br_2 is oxidised; Au is reduced.
- **C** Au^{3+} is oxidised; Br^- is reduced.
- **D** Br^- is oxidised; Au^{3+} is reduced.

2 In this reaction:
- **A** Au is the oxidising agent; Br_2 is the reducing agent.
- **B** Br_2 is the oxidising agent; Au is the reducing agent.
- **C** Au^{3+} is the oxidising agent; Br^- is the reducing agent.
- **D** Br^- is the oxidising agent; Au^{3+} is the reducing agent.

3 Identify the oxidation number of copper in $Cu(NO_3)_2$.
- **A** +3
- **B** +2
- **C** +1
- **D** –1

4 Identify the reduction half-equation.
- **A** $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
- **B** $Mn^{2+}(aq) + 4H_2O(l) \rightarrow MnO_4^{\ -}(aq) + 8H^+(aq) + 5e^-$
- **C** $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
- **D** $H_2(g) \rightarrow 2H^+(aq) + 2e^-$

5 A standard galvanic cell is set up with $Mg(s)/Mg^{2+}(aq)$ and $Cl_2(g)/Cl^-(aq)$ half-cells. In this cell:
- **A** Cl^- is the strongest oxidising agent and Mg^{2+} is the strongest reducing agent.
- **B** Cl_2 is the strongest oxidising agent and Mg is the strongest reducing agent.
- **C** Mg^{2+} is the strongest oxidising agent and Cl^- is the strongest reducing agent.
- **D** Mg is the strongest oxidising agent and Cl_2 is the strongest reducing agent.

6 A methane–oxygen fuel cell operates with the following half-equation at the anode:

$CH_4(g) + 2H_2O(l) \rightarrow CO_2(g) + 8H^+(aq) + 8e^-$

The reaction occurring at the cathode is:
- **A** $2H_2O(l) + 4e^- \rightarrow 4H^+(aq) + O_2(g)$
- **B** $4H^+(aq) + O_2(g) \rightarrow 2H_2O(l)$
- **C** $2H_2O(l) \rightarrow 4H^+(aq) + O_2(g) + 4e^-$
- **D** $4H^+(aq) + O_2(g) + 4e^- \rightarrow 2H_2O(l)$

The following information relates to Questions 7 and 8.

A galvanic cell is constructed using the following half-cells:

Half-cell 1: silver electrode and a colourless solution of $AgNO_3$

Half-cell 2: copper electrode in a blue solution of $CuCl_2$

7 Identify which of the following is likely to occur.
- **A** Chlorine gas forms at the anode.
- **B** The silver electrode decreases in mass.
- **C** The blue solution becomes darker.
- **D** The clear solution becomes cloudy.

8 A suitable salt bridge solution would contain:
- **A** NaCl
- **B** $NaNO_3$
- **C** NaOH
- **D** NaBr

9 Three half-cells, A, B and C, were constructed from unknown solutions according to the diagrams in Figure 1.

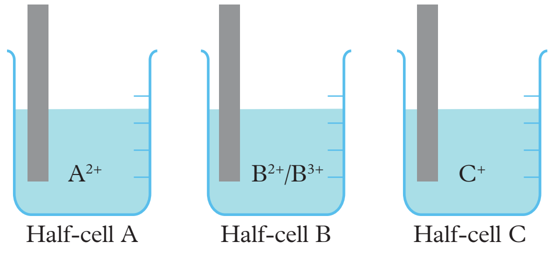


FIGURE 1 Three half-cells

The following information was determined experimentally when the half-cells were combined:

- The electrode in half-cell B is negative when combined with half-cell A.
- The electrode in half-cell C is negative when combined with half-cell B.

The strongest reducing agent is:

A A
B C^+
C B^{2+}
D C

10 Which of the following metals will not be oxidised by an aqueous 1.0 M acid solution?

A Copper
B Magnesium
C Zinc
D Iron

Short answer

Describe and explain

11 Describe the difference between the charge of an ion and its oxidation number.

12 Explain why corrosion is an example of oxidation.

13 Describe the main purposes of separating the half-cells in a galvanic cell.

14 Explain why galvanic cell voltages are measured at standard conditions.

15 Describe the purpose of an electrolyte, and how it is different from a salt bridge. Explain when an electrolyte is used.

16 Describe how the location of elements on the periodic table affects their ability to gain or lose electrons.

17 Explain the terms *oxidising agent* and *reducing agent*. Explain how you would use them to describe redox reactions.

Apply, analyse and compare

18 A student placed a piece of iron metal into a solution of lead(II) nitrate. Construct two half-equations to demonstrate the oxidation and reduction processes. Construct a balanced overall equation to represent the chemical reaction.

19 A piece of solid nickel was placed in a solution of blue copper(II) sulfate. Over time, the solution became paler in colour and an orange-coloured precipitate formed. Construct two half-equations to demonstrate the oxidation and reduction processes. Construct a balanced overall equation to represent the chemical reaction.

20 Determine the oxidation number of:

a N in NH_3
b Si in $CaSiO_3$
c Cr in $BaCrO_4$
d V in VO^{2+}.

21 Determine the oxidation number of iron in:

a Fe_2O_3
b FeO
c Fe_2SiO_4
d Fe.

22 Consider the following balanced chemical equation.

$$CO_2(g) + H_2(g) \rightarrow CO(g) + H_2O(g)$$

a Identify the oxidation numbers of all atoms.
b Determine which atom has been oxidised and which has been reduced.
c Determine which atoms are the oxidising and which are the reducing agents.
d Construct the oxidation and reduction half-equations.
e Construct the balanced overall redox reaction.

23 Consider the following balanced chemical equation.

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

a Identify the oxidation numbers for all atoms.
b Determine which atom has been oxidised and which has been reduced.
c Determine which atoms are the oxidising and which are the reducing agents.
d Construct the oxidation and reduction half-equations.
e Write a balanced overall redox reaction.

24 Consider the following unbalanced chemical equation.

$$Ag(s) + H_2S(g) \rightarrow Ag_2S(s) + H_2(g)$$

a Identify oxidation numbers of all atoms.
b Determine which atom has been oxidised and which has been reduced.

- **c** Determine which atoms are the oxidising and which are the reducing agents.
- **d** Construct the oxidation and reduction half-equations.
- **e** Write a balanced overall redox reaction.

25 Construct fully labelled galvanic cells for the following half-cells.

- **a** Ag/Ag^+ and Cu^{2+}/Cu
- **b** Cl_2/Cl^- and Fe^{2+}/Fe
- **c** H_2O_2/H^+ and O_2/H_2O_2

You must include:

- labelled anode, cathode, polarity of electrodes and the material the anode and cathode are made from
- fully balanced half-equations, labelled as oxidation or reduction (include states)
- a labelled salt bridge (propose a suitable chemical and demonstrate ion movement)
- the direction of electron flow
- all reactants and products in the two half-cells
- full calculations showing the E^0_{cell}
- the overall balanced chemical equation, including states.

26 A nickel–cadmium galvanic cell is constructed (Figure 2), where:

$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$ $\quad E^0 = -0.40$

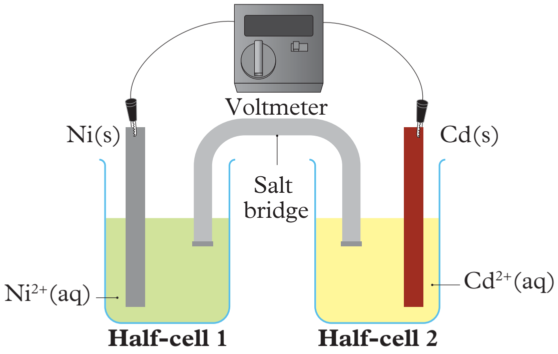


FIGURE 2 NiCd galvanic cell

Copy Figure 2 into your book and:

- **a** identify the oxidation and reduction half-cells
- **b** identify the anode and cathode, and the polarity of each electrode
- **c** construct the half-equations and overall equations
- **d** identify a suitable salt bridge substance and justify your choice
- **e** indicate the direction of ion and electron movement using appropriately labelled arrows
- **f** calculate the E^0 of the cell.

27 Consider the following pairs of half-cells:

- **a** Mg^{2+}/Mg and Fe^{2+}/Fe
- **b** Cu^{2+}/Cu and Ag^+/Ag
- **c** Al^{3+}/Al and H^+/H_2
- **d** Fe^{3+}/Fe^{2+} and $O_2 + H^+/H_2O_2$

For each pair of half-cells, state the:

- **i** reducing agent
- **ii** oxidising agent
- **iii** two half-equations
- **iv** overall equation.

28 An iron nail is placed in aqueous solutions of:

I $AgNO_3$
II $Mg(NO_3)_2$
III $CuSO_4$
IV $Pb(NO_3)_2$
V $Zn(NO_3)_2$

- **a** Identify the solution(s) in which you would expect the metal cation in the solution to plate onto the iron nail.
- **b** Use the electrochemical series to explain why some solutions would not result in the coating of the metal cation in the solution onto the nail.
- **c** Explain the purpose of the SO_4^{2-} and NO_3^- ions in the solutions.

29 Communicate the difference between galvanic cells, primary cells and fuel cells. Your answer must include:

- structural differences
- advantages
- disadvantages
- applications.

30 A copper–zinc galvanic cell operates for 15.0 minutes with a current of 0.25 A.

- **a** Write the half-equations of the cell and label them as oxidation and reduction.
- **b** Determine which electrode will lose mass and calculate the mass lost.

c Determine which electrode will gain mass and calculate the mass gained.

31 A galvanic cell is constructed using the following half-equations:

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l) \qquad E^0 = +1.36$

$Se(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2Se(g) \qquad E^0 = -0.40$

a Write the half-equations of the cell and label them as oxidation and reduction.

b The cell operates with a current of 2.5 A. Calculate the volume of gaseous H_2Se that is required to power the cell for 4.00 hours at SLC.

32 A hydrogen–oxygen fuel cell runs for 3.00 hours with a current of 5.0 A. Calculate the volume of oxygen gas required to operate the fuel cell at SLC.

Design and discuss

33 A student claimed that the best method to produce sodium hydroxide is via redox reactions between sodium metal and water to form sodium hydroxide and hydrogen gas. The half-equation for the water reaction is:

$2H_2O(l) + 2e^- \rightarrow 2OH^-(aq) + H_2(g)$

FIGURE 3 Hydrogen gas

a Write the oxidation and reduction half-equations.

b Construct a balanced overall equation to represent the chemical reaction.

c Evaluate the student's claim and discuss the strengths and weaknesses of this reaction.

34 Yttrium barium copper oxide ($YBa_2Cu_3O_7$) is a superconductor, meaning that it can conduct electricity very, very well. A claim was made that the oxidation numbers of all atoms in $YBa_2Cu_3O_7$ cannot be determined. Evaluate this claim and justify your opinion.

35 A student placed a piece of silver metal in a blue solution of copper(II) sulfate. After 10 minutes, the student observed that the solution had not changed colour. Evaluate the student's results and discuss a reason for this observation.

FIGURE 4 A piece of silver metal and a copper(II) sulfate solution

36 Discuss the advantages of attaching blocks of zinc to the steel supports, which contain iron, on oil rigs in Bass Strait.

37 A student wished to identify an unknown metal, *X*. The metal is cleaned, cut into 1 cm × 1 cm pieces and placed in separate solutions of:

- $Mg(NO_3)_2$
- $AgNO_3$
- $Al(NO_3)_3$
- $AuNO_3$.

The student observed that a layer of metal formed on metal *X* when placed in the silver and gold solutions, but observed no reaction when it was placed in the magnesium and aluminium solutions.

a Evaluate the results of the experiment and communicate what is happening to cause these results.

b Determine a possible identity for the metal and justify your choice using the electrochemical series.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

UNIT 3

AREA OF STUDY 1

Checkpoint

Multiple choice

Question 1

Which of the following fuels is renewable?

A Petrol
B Diesel
C Biodiesel
D Natural gas

Question 2

Which of the following provides the best definition of a renewable resource?

A A fuel that is harvested from crude oil and coal
B A fuel that can be replenished
C A fuel that can be replenished over several hundred thousand years
D A fuel that can be replenished within a short period of time

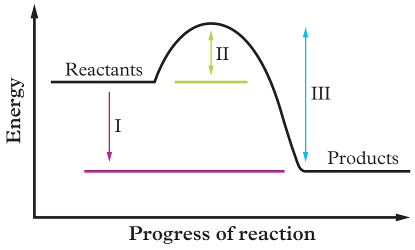
Question 3

Which of the following statements regarding renewable fuels is incorrect?

A They release greenhouse gases.
B They are derived from plant material or other organic sources.
C They have a higher energy density than non-renewable fuels.
D They can be replenished within the span of a lifetime.

Question 4

Which of the following quantities, I, II and III, are affected by the addition of a catalyst?



A II only
B II and III
C I only
D I, II and III

Question 5

The amount of energy released when a 20.0 L cylinder of ethane is combusted, according to the following equation, at SLC is:

$$2C_2H_6(g) + 14O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) \quad \Delta H = -3120\,\text{kJ mol}^{-1}$$

A 2516 kJ
B 1258 kJ
C 629 kJ
D 5032 kJ

Question 6

A butane camping cooker has a mass of 230 g, including the fuel. It is used four times a day over a period of five days. On each occasion, the fuel is used to heat a kettle containing 500 g of water from 4°C to 80°C. Calculate the total mass of the butane camping cooker after five days.

A 214.0 g
B 196.8 g
C 166.1 g
D 63.92 g

Question 7

The redox reaction between $Na_2Cr_2O_7$ and CH_3CH_2OH is given by the following chemical equation:

$$2Cr_2O_7^{2-}(aq) + 3CH_3CH_2OH(aq) + 16H^+(aq) \rightarrow 4Cr^{3+}(aq) + 3CH_3COOH(aq) + 11H_2O(l)$$

The correct half-equation for the oxidation process is:

A $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

B $CH_3CH_2OH(aq) + H_2O(l) \rightarrow CH_3COOH(aq) + H^+(aq) + e^-$

C $Cr_2O_7^{\ 2-}(aq) + 14H^+(aq) + 9e^- \rightarrow Cr^{3+}(aq) + 7H_2O(l)$

D $CH_3CH_2OH(aq) + H_2O(l) \rightarrow CH_3COOH(aq) + 4H^+(aq) + 4e^-$

Question 8

A galvanic cell is constructed using an iron(II) and tin half-cell. The tin half-cell contains a solution of $Sn(NO_3)_2(aq)$ and a solid tin electrode. The iron half-cell contains a mixture of $Fe(NO_3)_3$ and $Fe(NO_3)_2$, with a solid iron electrode. When the cell was built, the voltmeter was not recording an electrical current. A reason for this is:

A Sn is a stronger reducing agent than Fe, meaning that it will react in preference to it.

B Fe is the strongest reducing agent and will react spontaneously with Fe^{3+} ions, which are the strongest oxidising agents.

C Fe^{2+} cannot form Fe(s) and Fe^{3+} simultaneously.

D Fe is the strongest oxidising agent and will react spontaneously with Fe^{3+} ions, which are the strongest reducing agents.

Question 9

In an exothermic reaction:

A energy is absorbed from the surrounding environment.

B the energy of the products is greater than the energy of the reactants.

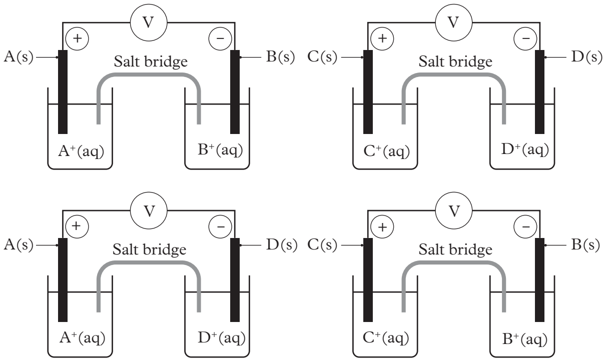
C the addition of a catalyst will increase the enthalpy of the reaction.

D the addition of a catalyst will decrease the activation energy.

Question 10

The following diagrams represent combinations of four half-cells (A/A^+, B/B^+, C/C^+ and D/D^+), which are investigated under standard conditions. Each half-cell consists of a metal electrode in a solution of 1.0 M metal nitrate. The polarity of the electrodes of each half-cell was determined and is represented in the circles next to the relevant electrode.

Which of the following represents the order of the metals from the strongest reducing agent to the weakest?



A B, A, C, D

B B, D, C, A

C D, A, C. B

D D, C, A, B

Short answer

Question 1 (4 marks)

Identical vehicle models were tested for fuel efficiency using petrol, E10 (petrol with 10% ethanol added), petrodiesel and biodiesel. Carbon dioxide emissions per litre of fuel burnt were also determined.

Fuel	Formula	Enthalpy ($kJ\,g^{-1}$)	Fuel consumption (L/100 km)	CO_2 produced (L of fuel)
Petrol	C_8H_{18}	48	14.5	2392
E10	90% C_8H_{18} 10% C_2H_5OH	44	14.2	2304
Petrodiesel	$C_{12}H_{26}$	44.8	9.2	3170
Biodiesel	$C_{19}H_{32}O_2$	42.0	7.5	2520

a Identify the fuel with the largest impact on the environment, in terms of the mass of CO_2 produced per 100 km travelled. Justify your answer with a calculation. 2 marks

b Describe one positive environmental impact of adding ethanol to petrol to form E10. 1 mark

c Explain why biodiesel is considered a renewable energy source but petrodiesel is not. 1 mark

Question 2 (9 marks)

Photosynthesis is represented by the following unbalanced equation:

$$CO_2(g) + H_2O(l) \rightarrow C_6H_{12}O_6(aq) + O_2(g)$$

The activation energy of the pathway catalysed by light is 2878 kJ mol^{-1}. The energy of the reactants is determined to be 1273.3 kJ mol^{-1}, and the energy of the products is 4087.3 kJ mol^{-1}.

a Determine the enthalpy for the photosynthesis reaction. 1 mark

b Construct a labelled enthalpy diagram for the process of cellular respiration. 2 marks

The glucose produced from photosynthesis can be fermented to produce bioethanol. This exothermic reaction has an enthalpy of –72 kJ mol^{-1}.

c Write a balanced thermochemical equation for the fermentation of glucose. 2 marks

d A 40.0 L fuel tank contains octane as a fuel in a car engine. If bioethanol is proposed as an alternative fuel to octane, calculate the volume of ethanol, and therefore the size of the fuel tank, that will produce as much energy as a 40.0 L tank of octane (density of octane = 0.703 g mL^{-1}; density of ethanol = 0.789 g mL^{-1}). 4 marks

Question 3 (8 marks)

To determine the molar enthalpy of dissociation of potassium nitrate (KNO_3), a student first calibrates a calorimeter by adding 100 g of water into the device. She then runs a current of 6.00 V and 1.20 A through the calorimeter for 5.50 minutes. The temperature of the water is found to increase from 12.0°C to 73.0°C.

a Calculate the calibration factor of the calorimeter. 2 marks

The calorimeter is emptied, refilled with 100 g of water and allowed to reach a constant temperature of 18.00°C. 2.500 g of solid potassium nitrate is then added to the calorimeter, with constant stirring. The final temperature of the solution is found to be 36.60°C.

b Calculate the molar enthalpy of the solution. 3 marks

c Compare the experimental value calculated to the theoretical molar enthalpy of –34.89 kJ mol^{-1}. 2 marks

d Describe one error which resulted in the difference in these values. 1 mark

Question 4 (5 marks)

In an experiment, half-cells are constructed and connected, one at a time, to a hydrogen reference half-cell. Although standard conditions were recommended, the temperature of the laboratory could not be controlled, but all solutions were made to a concentration of 1 M.

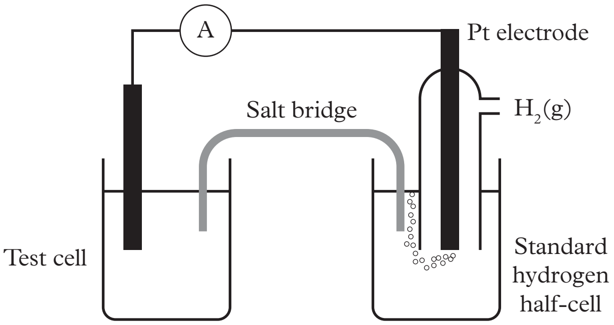


TABLE 1 Experimental results

Electrode material	Half-cell reaction	Observations
Pt	$SO_4{}^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons SO_2(aq) + 2H_2O(l)$	pH decreases in the standard hydrogen half-cell.
Ni	$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	Gas bubbles are observed at the cathode.
Pt	$MnO_4{}^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	pH increases in the test cell.

a Identify the strongest reducing agent from the half-equations in the table. 1 mark

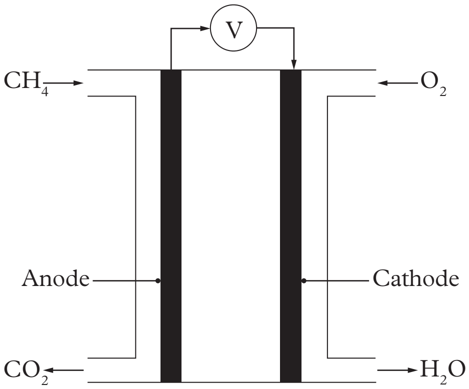
b Determine in which half-cell(s) the electrons flow from the test cell to the standard hydrogen half-cell. 1 mark

c State one flaw in the experiment that prevents the development of an electrochemical series. 1 mark

d Explain one change that can be made to the method to correct the flaw that you identified in part **c**. 2 marks

Question 5 (6 marks)

The fuel cell in the diagram below is constructed using an acidic electrolyte and natural gas as a fuel source. The electrodes are coated in platinum to produce electricity at low temperatures.



a State the polarity of the cathode. 1 mark

b Show the direction of the flow of ions in the electrolyte by drawing an arrow (→ or ←) in the space between the anode and cathode on the diagram. 1 mark

c Write a half-equation for the reaction that occurs at the anode of the cell. 1 mark

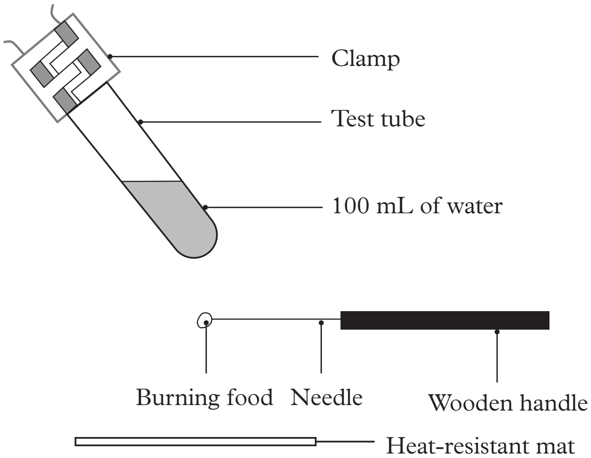
d Explain the purpose of the platinum in the cell. 2 marks

e Identify one way in which a fuel cell differs from a galvanic cell. 1 mark

Question 6 (8 marks)

To determine the heat of combustion or energy released (in kJ/100 g) from Cheezels and marshmallows, a food chemist places 100 g of water in a boiling test tube and allows it to reach room temperature before recording this as the initial temperature. They then weigh the food products and record these as their initial masses.

The food is skewered on a piece of wire and set alight under the test tube using a match. The final temperature of the water is recorded and the food products are reweighed to determine their mass after combustion.



The results of the experiment are shown in the table below.

a Calculate the enthalpy or heat of combustion (kJ/100 g) of the Cheezels. 2 marks

b Calculate the enthalpy or heat of combustion (kJ/100 g) of the marshmallow. 2 marks

c Compare the theoretical and experimental enthalpy values. 2 marks

d Comment on the accuracy and precision of the experiment and justify your responses. 2 marks

TOTAL MARKS

/50 marks

Food source	Initial mass (g)	Final mass (g)	Initial temperature of water (°C)	Final temperature of water (°C)	Enthalpy (kJ/100 g)
Cheezels	1.82	0.73	23.1	42.5	2270
Marshmallow	3.89	3.13	22.3	34.6	1356

CHAPTER 5

Rates of chemical reactions

KEY KNOWLEDGE

→ factors affecting the frequency and success of reactant particle collisions and rate of chemical reaction in open and closed systems including temperature, surface area, concentration, gas pressures, presence of a catalyst, activation energy and orientation

→ the role of catalysts in increasing the rate of specific reactions with reference to alternative reaction pathways of lower activation energies, represented using energy profile diagrams

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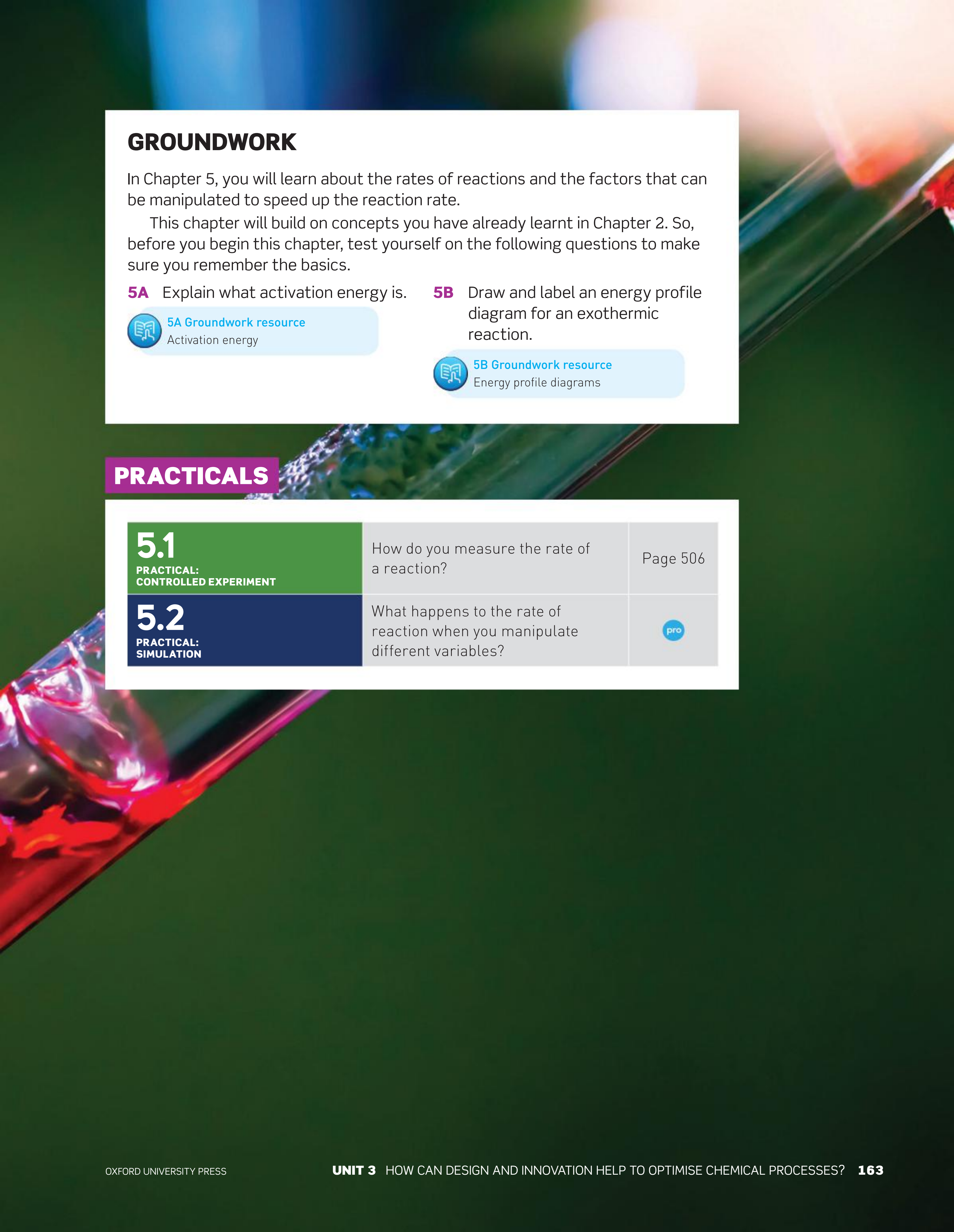
FIGURE 1 Heating reactants can increase the rate of reaction.



GROUNDWORK

In Chapter 5, you will learn about the rates of reactions and the factors that can be manipulated to speed up the reaction rate.

This chapter will build on concepts you have already learnt in Chapter 2. So, before you begin this chapter, test yourself on the following questions to make sure you remember the basics.

5A Explain what activation energy is.

5A Groundwork resource
Activation energy

5B Draw and label an energy profile diagram for an exothermic reaction.

5B Groundwork resource
Energy profile diagrams

PRACTICALS

5.1 **PRACTICAL: CONTROLLED EXPERIMENT**	How do you measure the rate of a reaction?	Page 506
5.2 **PRACTICAL: SIMULATION**	What happens to the rate of reaction when you manipulate different variables?	pro

5.1

Factors affecting chemical reactions

KEY IDEAS

In this topic, you will learn that:

- a reaction occurs when reactants collide successfully (with sufficient energy and in the correct orientation)
- a reaction rate can be increased by increasing the energy of a system, decreasing the activation energy, or increasing the frequency of collisions and therefore number of successful collisions
- concentration, pressure and surface area can affect the frequency of collisions and *number* of successful collisions
- only temperature will affect the frequency of collisions and number and *proportion* of successful collisions.

Scientists can manipulate chemical reactions to speed them up or slow them down. For example, industrial chemists work to strict deadlines and will often speed up chemical reactions to produce products quickly. Alternatively, food chemists will often slow down the rate of decay of nutrients and molecules within products to increase the supermarket shelf life of food. Manipulating the speed at which a reaction occurs requires an understanding of **reaction rates** (the change in the amount of products over time) and the factors that affect the success of a reaction.

reaction rate
the change in concentration of a reactant or product per unit of time (change over time)

Collision theory

Collision theory states that reactants must collide in order to react. However, not all collisions result in a reaction. In fact, only a certain percentage will collide successfully to form products. It is very common for reactants to collide with one another and bounce off without having reacted. For this reason, it is essential to understand the factors that result in a **successful collision**, in which reactants collide and react to form products.

collision theory
the theory that states that reactants must collide in the correct orientation and with sufficient energy for a reaction to occur

successful collision
when reactants collide with sufficient energy and in the correct orientation to form products

FIGURE 1 Food chemists manipulate reactions to slow the rate of decay and improve the shelf life of food.

These factors are as follows:

- Energy – Reactants must collide with a minimum amount of energy, called **activation energy**, to rearrange and form products. If C_2H_4 and HCl in Figure 2 do not collide with sufficient energy, the collision will not be successful.
- Orientation – Reactants must collide in the correct spot or in the correct position. Figure 2 demonstrates two molecules of C_2H_4 and HCl reacting together according to the chemical equation:

  $$C_2H_4(g) + HCl(g) \rightarrow C_2H_5Cl(g)$$

  The C_2H_4 and HCl must collide in the correct orientation, as shown in collision 1; otherwise, the collision will be unsuccessful.

activation energy
the minimum amount of energy that a collision must have for a reaction to occur

5.1
PRACTICAL: CONTROLLED EXPERIMENT

How do you measure the rate of a reaction? Go to page **506**.

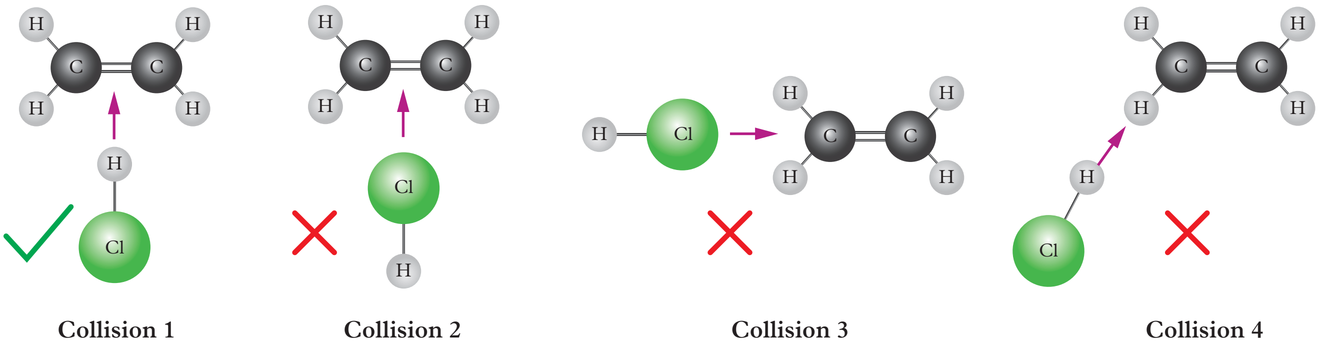


FIGURE 2 The correct and incorrect orientation of ethene (C_2H_4) and hydrochloric acid (HCl) molecules for a successful collision. Collision 2 is not successful because of the repulsion between the slightly negatively charged chlorine atom and the electrons in ethene's double bond.

Study tip

Collisions can occur without resulting in a chemical reaction, so it is essential to use the term *successful collision* when explaining whether a chemical reaction results in the formation of a product.

Activation energy on energy profile diagrams

The law of conservation of energy states that energy cannot be created or destroyed. Chemical energy is a type of potential energy, because it is the energy stored within the bonds of reactants and products. This chemical energy can be readily converted into heat energy as temperature or motion (kinetic energy). Energy differences between reactants and products result in either an **exothermic** reaction, where excess energy is released into the environment, or an **endothermic** reaction, where energy is absorbed from the environment.

Energy profile diagrams represent the energy changes that occur in chemical reactions. These diagrams show the difference in energy within the bonds of reactants and products. Activation energy, the minimum energy needed for a reaction to occur successfully, is represented as E_a in Figure 3. All reactants must overcome this energy to break reactant bonds when they collide, and therefore create a successful reaction.

exothermic
if the enthalpy of the products is lower than the enthalpy of the reactants, the reaction loses energy as heat to the environment

endothermic
if the enthalpy of the reactants is lower than the enthalpy of the products, the reaction must absorb energy as heat from the environment

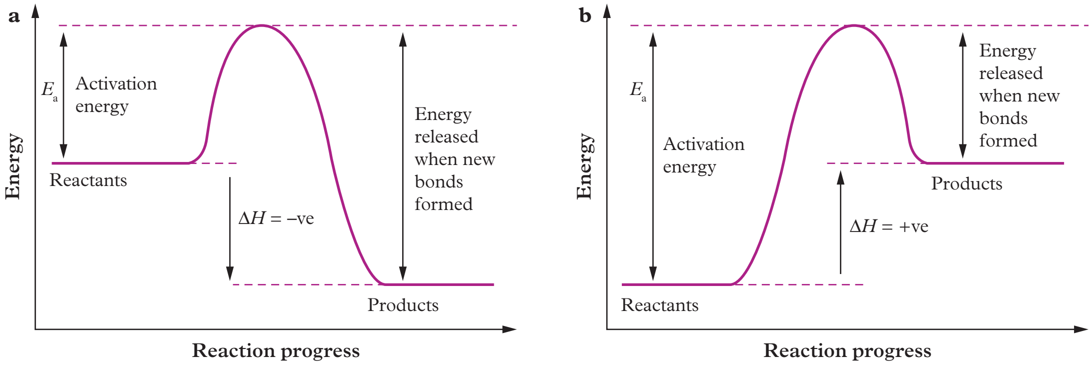


FIGURE 3 Energy profile diagrams for **a** exothermic and **b** endothermic reactions, including E_a (the activation energy required to break bonds) and ΔH (enthalpy)

Study tip

Think of the terms *exothermic* and *endothermic* like a party. *Exo* means exit, when friends leave the party – it's bad, so it is negative. *Endo* means enter, when friends arrive at the party – it's great, so it is positive.

Magnitude of activation energy

The reaction rate depends on the size of the activation energy. If the energy contained within the bonds of the reactants is relatively low (i.e. the bonds are weak), lower activation energy is required. Conversely, strong bonds contain higher amounts of energy and are harder to break apart, so they require higher activation energy.

Bond energy is the amount of energy stored within a bond. The higher the energy contained within the bond, the harder it is to break. Therefore, bond energy (strength) and the number of bonds also affect the activation energy and how readily a reaction occurs.

bond energy
a measure of the strength of a chemical bond, in kJ mol^{-1}

Maxwell–Boltzmann energy distribution

The particles within a reaction system do not all contain the same amount of energy. Some have more energy than others. The Maxwell–Boltzmann distribution (Figure 4) represents the number of particles in a reaction system (on the y-axis) and the kinetic energy, E, of each particle (on the x-axis).

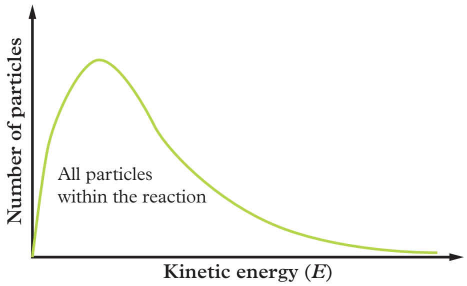


FIGURE 4 A Maxwell–Boltzmann energy distribution curve

The peak of the curve represents the average kinetic energy of the particles and is relative to the temperature of the reaction. The area underneath the curve represents all particles within the reaction.

A Maxwell–Boltzmann energy distribution is usually plotted with a line representing the activation energy (E_a) of the reaction. Only the particles that have kinetic energy greater than or equal to the activation energy ($E \geq E_a$) will react successfully when they collide.

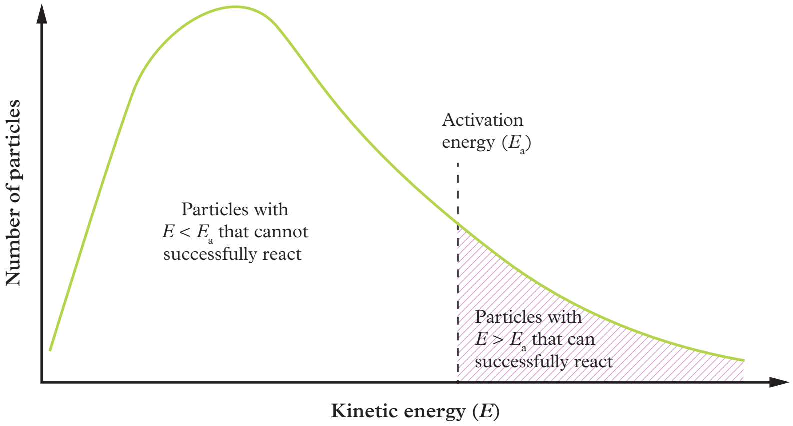


FIGURE 5 The Maxwell–Boltzmann distribution demonstrates the energy of particles in a reaction system.

Study tip

The peak of a Maxwell–Boltzmann distribution is the average kinetic energy of particles, which represents the temperature of the reaction.

number of successful collisions
the total number of reactant particles in the reaction that are colliding successfully and forming products

proportion of successful collisions
the proportion or percentage of reactant particles in the reaction that are colliding successfully and forming products

frequency of collisions
how often collisions occur between reactants (including successful and unsuccessful collisions)

Factors affecting the frequency of collisions

A reaction can take anywhere from a few hours to more than a few hundred thousand years, so it is important that scientists are able to increase the rate of a chemical reaction (to make it happen faster). The reaction rate can be increased by:

- increasing the energy of a reaction
- decreasing the activation energy
- increasing the frequency, and therefore **number or proportion of successful collisions**.

Increasing the number of successful collisions can be accomplished by manipulating five main factors, depending on the state of the reactants and products. These factors are:

- concentration: liquid, aqueous and gaseous reactants
- pressure: gaseous reactants only
- surface area: solids reactants only
- temperature: solids, liquids, aqueous and gaseous reactants
- catalysts: solids, liquids, aqueous and gaseous reactants (explored in Topic 5.2).

When explaining how each factor affects the rate of a chemical reaction, you must always discuss the **frequency of collisions**: how often a collision occurs and the number or proportion of successful collisions.

Concentration

Concentration is the measure of the number of particles in a given volume. Concentration can be applied to solid, liquid, aqueous and gaseous systems. An increase in concentration means an increase in the number of particles per unit of volume. If the number of particles increases, the particles are closer together. This increases the chance of reactants colliding and therefore increases the frequency of collisions. If collisions occur more frequently, there will be a higher number of successful collisions and the rate of reaction will increase.

concentration
a measure of the number of particles in a volume of substance

Alternatively, dilute solutions have fewer reactant particles, resulting in fewer collisions. A decrease in reactant particles decreases the frequency of collisions and the number of successful collisions, and therefore the rate of reaction.

In aqueous solutions, concentration refers to the amount of solute in a volume of solvent. In gases, it is the number of gas molecules in a measured container volume. In a solid, the concentration is the number of particles in a larger volume of a solid substance (or its density).

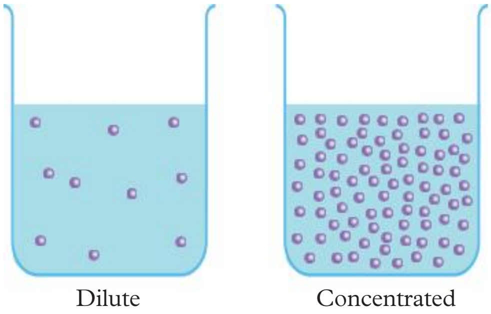


FIGURE 6 A dilute solution has fewer particles than a concentrated solution, so it results in fewer collisions.

Study tip

An increase in concentration will increase the frequency of collisions and therefore the number of successful collisions.

Pressure

Pressure is the force exerted per unit area by one substance upon another substance. For gases, this is the pressure exerted by gas particles when they collide with the walls of a container. An increase in the number of particles in a system results in an increase of pressure.

pressure
the force exerted, per unit area, by one substance upon another substance

A gaseous reaction is often referred to as a system, because it must be enclosed within a container or the gases will escape. Increasing the volume of the system is the equivalent of increasing the volume of the container that is holding the gaseous reaction.

The pressure of a gaseous system can be increased in two ways:

1 Decreasing the volume of the container – this pushes the particles closer together and results in a greater number of particles per unit of volume (Figure 8). For example, the concentration of 1000 oxygen molecules in a 1 L container is lower than that of 1000 oxygen molecules in a 100 mL container.

   In the 100 mL container, the particles are closer together and collide more frequently, resulting in a higher number of successful collisions.

2 Increasing the number of particles within the container – in a fixed-volume system, the volume cannot be decreased. Therefore, to increase the pressure, more gas particles must be added to the container, which increases the number and concentration of particles (Figure 8).

   An increase in the number of particles will cause particles to collide more frequently, resulting in a higher number of successful collisions.

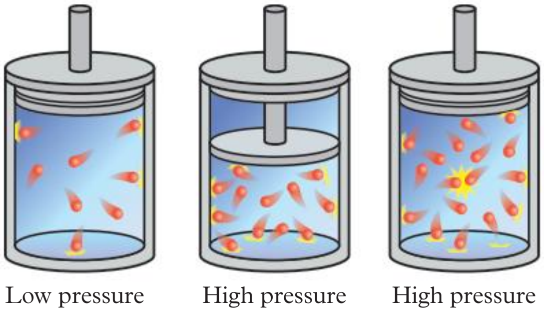
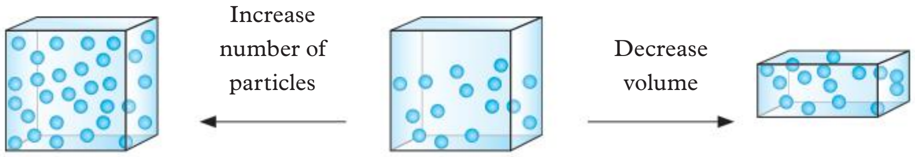


FIGURE 7 Visual representations of low pressure and high-pressure gaseous systems



n	Original gas container with 30 gas molecules	Original gas container with 15 gas molecules	Half of the original container with 15 gas molecules
V	1 L	1 L	500 mL or 0.5 L
c	$c = \frac{n}{V}$ $= \frac{30}{1}$ = 30 particles per litre	$c = \frac{n}{V}$ $= \frac{15}{1}$ = 15 particles per litre	$c = \frac{n}{V}$ $= \frac{15}{0.5}$ = 30 particles per litre

FIGURE 8 The effect on concentration when pressure is increased by increasing the number of particles or decreasing the volume

In both methods, an increase in pressure is due to an increase in the concentration of particles. This causes particles to collide more frequently and results in a greater number of successful collisions, which increases the rate of reaction.

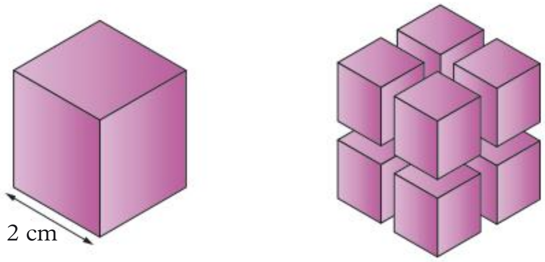
Study tip

An increase in the surface area will increase the frequency of collisions and therefore the number of successful collisions and the rate of reaction.

Surface area

Reactions occur at the surface of a solid reactant. Therefore, the greater the **surface area**, the greater the number of collisions. Figure 9 shows a 2 cm × 2 cm cube with a surface area of 24 cm^2. When the cube is cut into eight smaller 1 cm × 1 cm cubes, the total surface area increases to 48 cm^2. Therefore, a solid that has been broken down into powder has a larger surface area than a larger solid crystal. This increases the frequency of collisions in a reaction, the number of successful collisions, and therefore the rate of reaction.

surface area
a measure of the total surface area available to react in a chemical reaction



Total surface area
= 2 cm × 2 cm × 6 sides
= 24 cm^2

Total surface area
= 1 cm × 1 cm × 6 sides × 8 cubes
= 48 cm^2

FIGURE 9 The surface area calculation for a bigger 2 cm × 2 cm cube and smaller 1 cm × 1 cm cubes

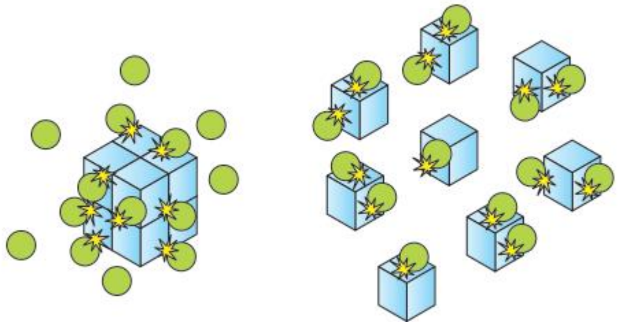
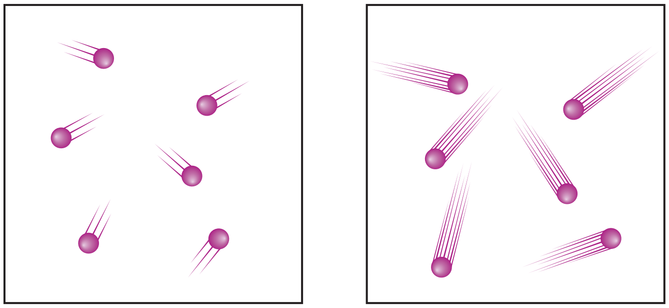


FIGURE 10 Increasing the surface area of a reactant by breaking it into smaller pieces increases the number of successful collisions and therefore the reaction rate.

Temperature

Kinetic molecular theory states that all particles are in constant, random motion. The energy involved in this motion is kinetic energy. Heating particles increases their movement and therefore their kinetic energy. Particles collide more frequently, increasing the number of successful collisions. Additionally, a greater proportion of particles will have energy greater than, or equal to, the activation energy. This increases the *proportion* of successful collisions. Together, this increases the rate of reaction.

If a substance is cold, its particles will have lower kinetic energy, the frequency of collisions, and number and proportion of successful collisions will decrease. This decreases the rate of reaction. (Figure 11).



Cold substances (particles have low kinetic energy)

Hot substances (particles have high kinetic energy)

FIGURE 11 Increased temperature means increased kinetic energy of particles, resulting in faster movement.

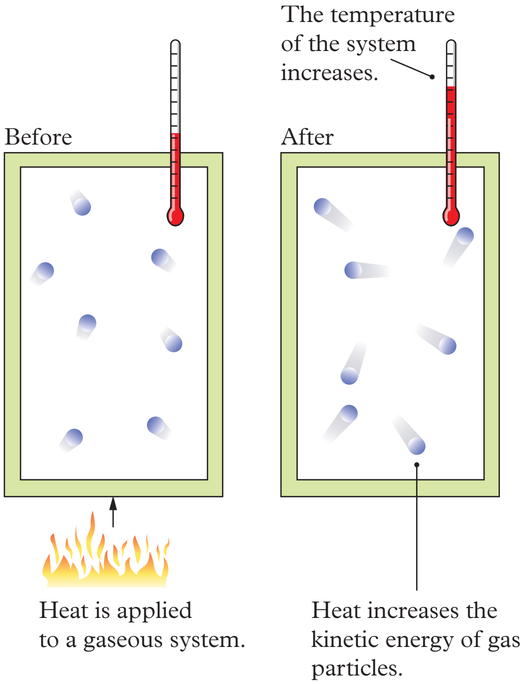


FIGURE 12 Heating a reactant increases the kinetic energy of particles.

The Maxwell-Boltzmann curve

The peak of a Maxwell–Boltzmann curve is the kinetic energy of the highest number of particles. This peak represents the average kinetic energy of the reaction and also equates to the temperature of the system (Figure 13).

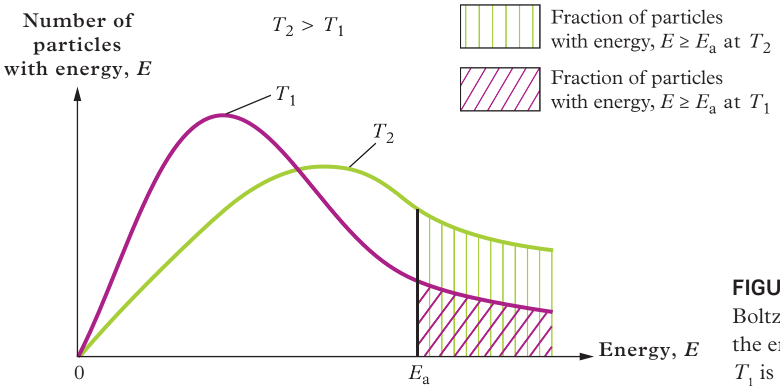


FIGURE 13 The Maxwell–Boltzmann distribution demonstrates the energy of particles in a system. T_1 is a lower temperature than T_2.

Several key pieces of information can be gained from this distribution.

- The peak of the distribution represents the average kinetic energy; the further this peak is to the right, the higher the average temperature of the system.
- The areas under both curves must be equal as they represent the same reaction, with an identical number of particles, but at different temperatures.

kinetic molecular theory
the theory that states that all particles are in constant random motion

Study tip

Temperature changes the *proportion* of successful collisions, which has a more significant effect on reaction rate than the *number* of successful collisions.

Study tip

When drawing reaction curves on a Maxwell–Boltzmann distribution:

- at a higher temperature, the peak of the curve is shifted right and the curve is lower and broader
- at a lower temperature, the peak of the curve is shifted left and the curve is higher and narrower.

- At higher temperatures, the peak of the curve shifts to the right; the curve is both flatter and broader than the lower-temperature curve.
- At lower temperatures, the peak of the curve shifts to the left; the curve is both higher and narrower than the higher-temperature curve.
- At higher temperatures, there is a higher proportion of particles with higher kinetic energy. Therefore, there are more particles with energy greater than the activation energy ($E > E_a$).

In Figure 13, the distribution clearly demonstrates that at the higher temperature (T_2), there are more particles with energy greater than the activation energy ($E > E_a$). This results in more successful collisions and a faster reaction.

Study tip

An increase in temperature will increase the frequency of collisions with energy greater than or equal to the activation energy ($E \geq E_a$). This will increase the proportion of successful collisions and the rate of reaction.

5.1 Worked example
Find me in your obook pro

5.1 Worked example
Video demonstration

5.1 Challenge
Find me in your obook pro

5.1 REAL-WORLD CHEMISTRY

The problem with dust – the sugar refinery explosion

At 7 pm on 7 February 2008, an explosion rang out across the town of Port Wentworth, Georgia, USA. The blast originated in the centre of the Georgia Sugar Refinery, located on the banks of the Savannah River. The refinery, built in 1916, was old. Its construction features were outdated, and many staff members said the machinery was antiquated, some of it being close to 30 years old.

The explosion took place in the centre of the refinery where the sugar was stored before being bagged and packaged.

Refined sugar is sucrose, a disaccharide, which consists of one glucose monomer and one fructose monomer. When bonded together, the disaccharide has the molecular formula $C_{12}H_{22}O_{11}$. Sucrose undergoes combustion according to the chemical equation:

$$C_{12}H_{22}O_{11}(s) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(g) \qquad \Delta H = -5645\,\text{kJ}\,\text{mol}^{-1}$$

Sugar is naturally highly combustible; however, as a fine dust powder, it can also spontaneously combust. The storage of the sugar led to a build-up of sugar dust within the air. The building was equipped with ventilation fans that extracted dust particles from the air, but these were found to contain enough dust to cause further explosions. The ignition source was probably static electricity or a spark, but spontaneous combustion was not ruled out.

The fire took seven days to extinguish because of the age of the building and the quantity of sugar that continued to act as a fuel source. The task of putting out the fire was made worse because the sugar that did not combust melted at temperatures above 186°C.

The explosion killed 14 people and severely injured 36.

Apply your understanding

1. With reference to the frequency and proportion of successful collisions, explain how powdered sugar can explode, whereas sugar cubes are relatively safe.
2. Using your knowledge of the rate of reaction, explain the recommendations you would make for future plants in terms of the storage of sugar. Hint: Consider the reactants in your answer.
3. Discuss the safety precautions the refinery should have taken to minimise the impact of the dust.

Study tip

You will lose marks in your exam if you do not mention both the *frequency of collisions* and *number* or *proportion of successful collisions.*

Study tip

Remember that concentration, surface area and pressure deal with *frequency* and *number*. Temperature deals with *frequency, number* and *proportion.*

5.1 CHECK YOUR LEARNING

Describe and explain

1 Describe the factors that determine whether a collision is successful or not.

2 Describe the factors that decrease the rate of a chemical reaction.

3 Define the term *activation energy* and explain why it is an essential factor in the rate of a chemical reaction.

Apply, analyse and compare

4 Apply your knowledge of rates of reaction to the following chemical reaction:

$$CaCl_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + 2HCl(aq) \quad \Delta H = \text{negative}$$

Explain, with reference to the frequency and proportion of successful collisions, the factors that could increase the rate of this chemical reaction.

5 A chemist wants to optimise the rate of the chemical reaction below:

$$CO_2(g) + NO_2(g) \rightarrow CO_2(g) + NO(g)$$

a Explain, with reference to the frequency and number of successful collisions, what three factors can be investigated that increase the rate of this reaction.

b The Maxwell–Boltzmann distribution of particle energies at 30°C is shown below.

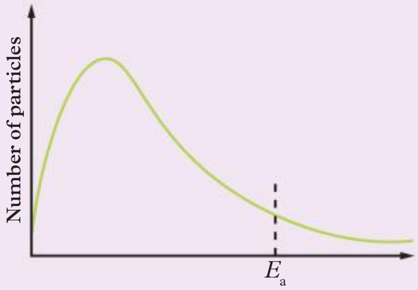


FIGURE 14 The Maxwell–Boltzmann energy distribution of the reaction between carbon monoxide and nitrogen dioxide

Sketch this diagram into your exercise book and, on the same diagram, sketch the curve that results when the temperature of the same reaction is:

i increased **ii** decreased.

c Explain the meaning of the peak of the three curves.

d Use these three curves to explain the effect of increasing temperature on the rate of reaction.

6 The following Maxwell–Boltzmann distribution was developed for a reaction mixture labelled Sample 1. A second sample of the same reaction mixture in the same reaction vessel is labelled as Sample 2.

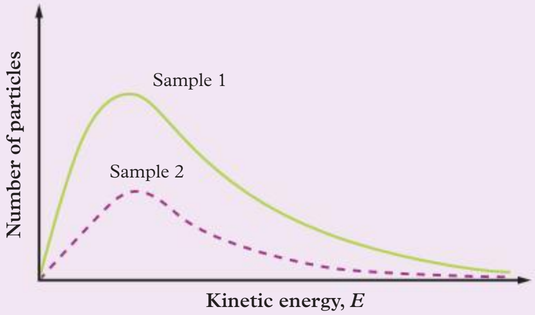


FIGURE 15 The Maxwell–Boltzmann energy distribution of the same reaction mixture when a change has been made

Analyse the curves in Figure 15 and explain what changes have been made to the reaction between Sample 1 and Sample 2 to result in the different Maxwell–Boltzmann curves. Explain which has a faster rate of reaction.

Design and discuss

7 Iron rusts according to the following equation:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

This equation occurs in the presence of water. A car manufacturer has found that their car parts are rusting before the painting and assembly processes. Discuss the recommendations that you would make to the manufacturer to slow the rusting process or avoid it entirely.

8 Sugar cubes were used many years ago as a sweetener for tea and coffee, but now granulated sugar is more commonly used. Discuss why this change was made – i.e. the benefits of using granulated over cubed sugar.

5.2

The role of catalysts

KEY IDEAS

In this topic, you will learn that:

- there are two types of catalysts: homogeneous catalysts, which are in the same state as reactants, and heterogeneous catalysts, which are in a different state from reactants
- catalysts increase the rate of reactions by providing an alternative reaction pathway with reduced activation energy.

catalyst
a substance that lowers the activation energy of a reaction and increases the rate of reaction without participating in the chemical reaction

homogeneous catalyst
a catalyst that is in the same state as the reactants

heterogeneous catalyst
a catalyst that is in a different state from the reactants

Study tip

Homogenous and homogeneous mean the same thing, as do heterogenous and heterogeneous. The terms that include an *e* are more modern versions, with the same meaning.

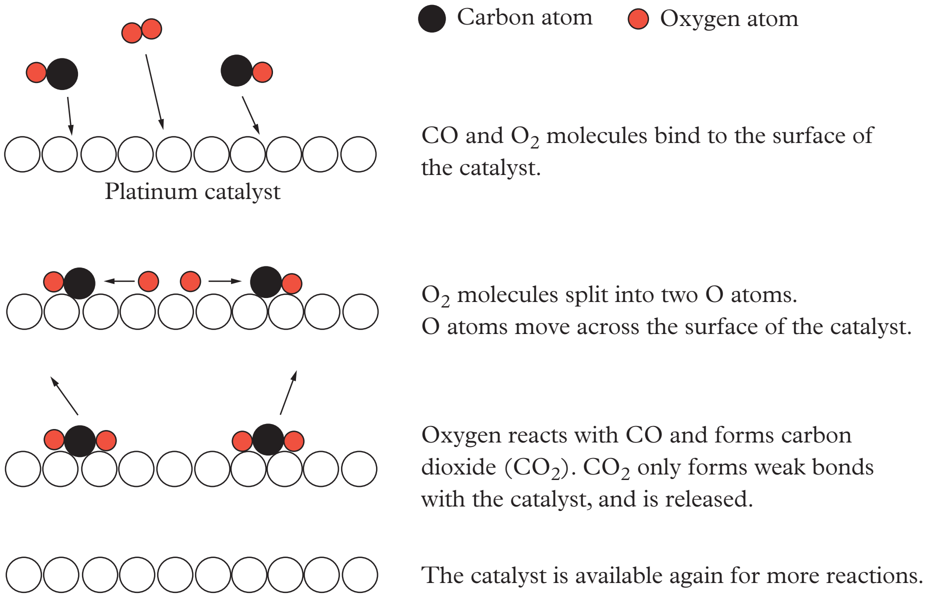
Catalysts are substances that speed up the rate of a chemical reaction without participating in the reaction. They do this by providing an alternative reaction pathway that requires a lower activation energy. Catalysts are essential in the human body, medicine, pharmaceuticals, the chemical industry and many other science-related fields.

There are two types of catalysts.

- **Homogeneous catalysts** are in the same state as the reactants. For example, in the dissociation of aqueous hydrogen peroxide ($H_2O_2(aq)$), an aqueous potassium iodide catalyst is used.
- **Heterogeneous catalysts** are in a different state from the reactants. For example, in the reaction between hydrogen and nitrogen, a solid iron catalyst is used.

How catalysts work

Catalysts affect the rate of reactions by providing an alternative reaction pathway with reduced activation energy. Although the mechanism of how catalysts lower activation energy is not thoroughly understood, scientists know that catalysts have high surface energy. When reactants strike the surface of the catalyst, they bond to it. The formation of these bonds weakens or breaks the bonds contained within the reactants. The atoms can then rearrange on the catalyst's surface to form products (Figure 1).



CO and O_2 molecules bind to the surface of the catalyst.

O_2 molecules split into two O atoms.
O atoms move across the surface of the catalyst.

Oxygen reacts with CO and forms carbon dioxide (CO_2). CO_2 only forms weak bonds with the catalyst, and is released.

The catalyst is available again for more reactions.

FIGURE 1 The action of a catalyst in lowering the energy required for a reaction to occur

5.2
PRACTICAL: SIMULATION

What happens to the rate of reaction when you manipulate different variables? Go to your **Student obook pro**.

The effect of a catalyst on the Maxwell–Boltzmann distribution

The catalysed reaction can be represented on a Maxwell–Boltzmann distribution as a line, similar to the one that represents activation energy. Because a catalyst lowers the activation energy of a system, the number of particles with energy greater than the activation energy ($E > E_a$) increases (Figure 2).

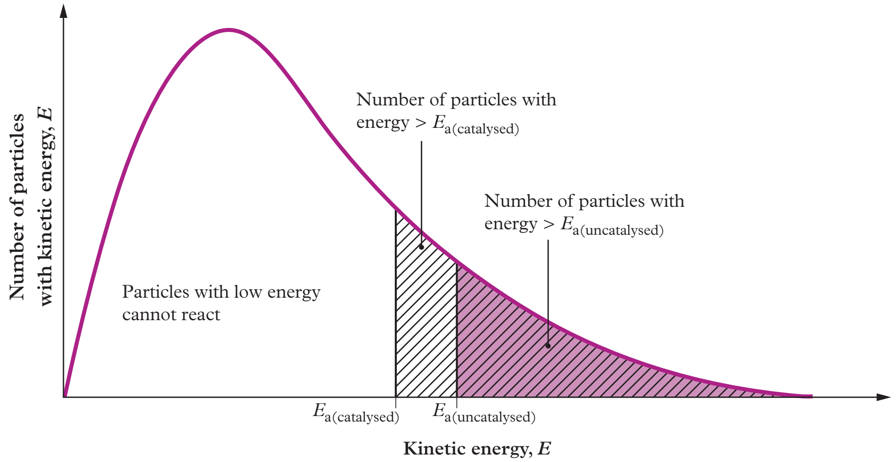


FIGURE 2 The Maxwell–Boltzmann distribution demonstrates how activation energy (E_a) is lowered with a catalyst.

The effect of a catalyst on an energy diagram

The way a catalyst lowers the E_a of a reaction can also be represented on an energy profile diagram. In Figure 3, the catalysed chemical pathway is represented as a broken line below the uncatalysed chemical pathway. The activation energy of the catalysed pathway (E_c) is lower than E_a, but the enthalpy of the reaction (ΔH) never changes.

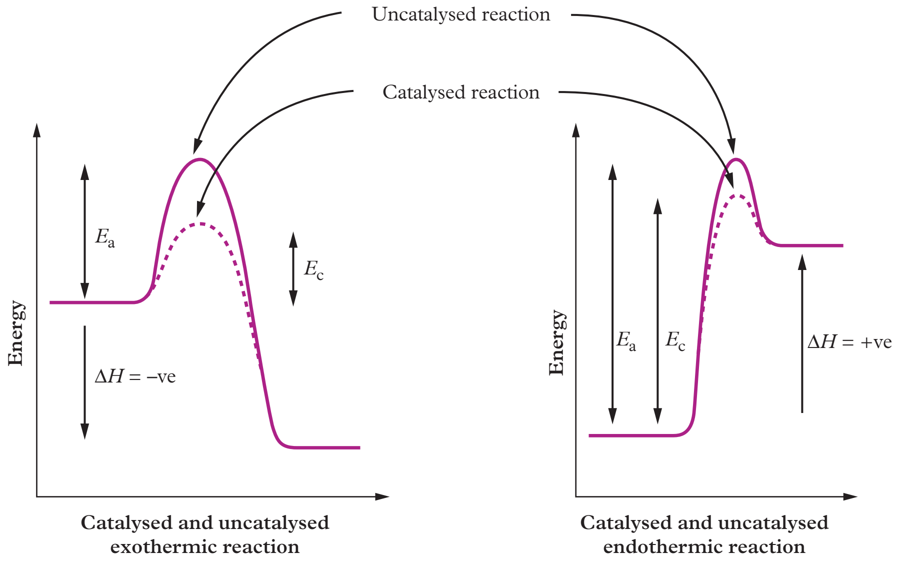


FIGURE 3 Energy profile diagrams of exothermic and endothermic reactions, demonstrating the regular reaction pathway and the catalysed reaction pathway with lowered activation energy. E_a = activation energy of the uncatalysed reaction; E_c = activation energy of the catalysed reaction.

The effect of a catalyst on the frequency and proportion of successful collisions

Catalysts do not increase the kinetic energy of the reactants, nor do they push particles closer together so that they collide more often. For this reason, catalysts do not increase the frequency of the collisions within a reaction. Instead, catalysts lower the activation energy, resulting in a greater proportion of particles within the reaction that have energy greater than or equal to the activation energy. Therefore, the use of a catalyst increases the proportion of successful collisions.

Study tip

Adding a catalyst will not increase the frequency of collisions. It causes a greater proportion of particles to have $E \geq E_a$, resulting in a greater proportion of successful collisions.

5.2 SKILL DRILL

Analysing and evaluating an experiment using a catalyst

Key science skill: Analyse and evaluate data and investigation methods

In an experiment to determine the effect of an increase in concentration on the rate of reaction, a student investigated the rate of decomposition of hydrogen peroxide into water and oxygen gas using a potassium iodide catalyst.

$$H_2O_2(aq) \rightarrow H_2O(l) + O_2(g) \quad \Delta H = -191.2\,\text{kJ}\,\text{mol}^{-1}$$

The student used 50.00 mL of room temperature (18°C) 1 M, 2 M and 3 M of $H_2O_2(aq)$ and measured the mass of the gas lost every 10 seconds by placing the reaction on a balance. Each concentration was repeated three times and the results were found to vary significantly between repeats.

Practise your skills

1 Identify the independent and dependent variables.
2 List as many controlled variables as possible given the information above.
3 If the reaction is exothermic, explain whether the controlled variables can be kept constant.
4 Comment on the precision, reliability and validity of the experiment.

Need help analysing and evaluating investigation methods? See Topic 1.8 (page 24).

5.2 CHALLENGE

Comparing the states of catalysts

An aqueous chemical reaction was conducted using either a homogeneous or a heterogeneous catalyst.

1 Identify the states of the homogeneous and heterogeneous catalysts in this reaction.
2 Identify which one resulted in a faster rate of reaction and explain your reasoning.
3 Explain your reasoning for your answer to Question 2.

Table 1 summarises the factors that affect the frequency, number and proportion of succcessful collisions.

TABLE 1 How the factors affect collisions

Factor	Affects…		
	Frequency of collisions	**Number of successful collisions**	**Proportion of successful collisions**
Concentration	✓	✓	
Surface area	✓	✓	
Pressure	✓	✓	
Temperature	✓	✓	✓
Catalyst		✓	✓

5.2 CHECK YOUR LEARNING

Describe and explain

1 Explain the effect a catalyst has on the rate of a chemical reaction.

2 Explain the difference between a homogeneous and a heterogeneous catalyst.

FIGURE 4 Different solid catalysts

3 Describe the effect that a catalyst has on the activation energy and enthalpy of the reaction.

Apply, analyse and compare

4 Draw a labelled Maxwell–Boltzmann diagram representing particles in a reaction at 30°C and 60°C. Draw and label the activation energy and the catalyst's activation energy.

5 The dissociation of hydrogen peroxide (H_2O_2), using a potassium iodide catalyst, is an exothermic process that forms water and oxygen gas. H_2O_2 contains $498\,kJ\,mol^{-1}$ of energy and the enthalpy of the reaction is $196\,kJ\,mol^{-1}$. The activation energy is approximately $75\,kJ\,mol^{-1}$, which is reduced by $26\,kJ\,mol^{-1}$ when the catalyst is added. Draw a fully labelled energy profile diagram of the dissociation of H_2O_2. Demonstrate both the catalysed and uncatalysed pathways.

Design and discuss

6 Investigate the production of ethanol by the fermentation of sugar and the reaction of water with ethene.

a Write a balanced chemical equation for each chemical reaction.

b Identify the catalyst involved in each chemical reaction.

c Comment on the environmental impact of each reaction.

7 $Fe(NO_3)_3(aq)$ and $MnO_2(s)$ have been proposed as alternative catalysts to KI in the dissociation of peroxide. 20.00 mL of 35%(v/v) $H_2O_2(aq)$, which has a density of $1.45\,g\,mL^{-1}$, was used in the reaction.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

FIGURE 5 Aqueous $Fe(NO_3)_3$ is a potential catalyst.

a Explain which catalyst you would predict to cause a faster rate of reaction.

b Identify the variables that must be controlled in the reaction.

c Explain what should be done to increase the validity of the experiment.

d Explain how the precision of the data can be increased.

e Calculate the theoretical mass loss of oxygen in the experiment.

f Construct a graph of the mass of oxygen generated over time.

g To the graph from part **f**, add a second trend line to represent the catalysed reaction.

CHAPTER 5 Review

Chapter summary

5.1

- Collision theory states that reactants must collide in the correct orientation and with sufficient energy to result in a successful chemical reaction.
- To increase the rate of reaction, the activation energy (minimum energy required for a chemical reaction to occur) must be decreased or the number of collisions resulting in a successful collision must be increased.
- An increase in either concentration, pressure (and therefore decrease in volume) or surface area will result in an increase in the frequency of collisions and the number of successful collisions.
- The Maxwell–Boltzmann distribution is used to represent the kinetic energy of all particles in a reaction, with the peak of the curve being the average kinetic energy and temperature. The presence of the E_a line indicates the number of particles in a reaction that have greater energy than the activation energy and will therefore successfully react.
- An increase in temperature will increase the frequency of collisions, and number and proportion of successful collisions.

5.2

- Catalysts provide an alternative pathway, which requires less activation energy, resulting in an increased rate of chemical reaction. Catalysts can be represented on the Maxwell–Boltzmann distribution as well as on an energy profile diagram.
- Catalysts do not increase the frequency of collisions, but they do increase the number and proportion of successful collisions.

Key formulas

Concentration	$n = c \times V$

Chapter checklist

Use the success criteria in the below to rate how well you understand each concept as ‘Confidently’, ‘Mostly’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Mostly	Not really	Revision link
• explain the concept of collision theory	☐	☐	☐	Go back to Topic 5.1
• describe how temperature, surface area, pressure and concentration can affect the rate of reaction	☐	☐	☐	Go back to Topic 5.1
• describe how a catalyst increases the rate of reaction	☐	☐	☐	Go back to Topic 5.2
• distinguish between catalysed and uncatalysed reaction pathways on Maxwell–Boltzmann distributions and energy profile diagrams	☐	☐	☐	Go back to Topic 5.2

Revision questions

Multiple choice

1 Consider the following chemical reaction:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Which of the following will not increase the rate of reaction?

A Crushing the magnesium into a powder

B Increasing the concentration of oxygen

C Adding water to make the reactants aqueous

D Increasing the temperature by burning the magnesium

2 Which of the following will not increase the rate of the following chemical reaction?

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

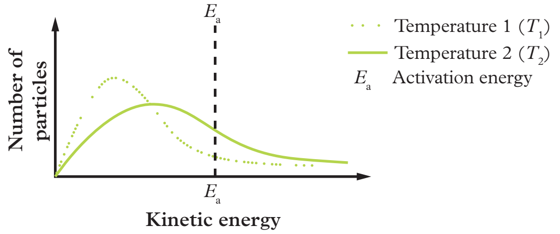
A Increasing the volume of the container

B Increasing the concentration of NaCl

C Increasing the temperature

D Increasing the concentration of $AgNO_3$

3 Which of the following conclusions about the following diagram is correct?



A At T_2, fewer particles have greater energy than E_a, so the rate of reaction will be slower.

B At T_1, fewer particles have greater energy than E_a, so the rate of reaction will be faster.

C At T_1, more particles have greater energy than E_a, so the rate of reaction will be slower.

D At T_2, more particles have greater energy than E_a, so the rate of reaction will be faster.

4 In a catalysed chemical reaction:

A the difference in enthalpy between products and reactants (ΔH) changes.

B an alternative reaction pathway with lower activation energy is provided.

C the kinetic energy of particles and therefore temperature increases.

D an alternative reaction pathway with higher activation energy is provided.

5 A catalyst increases the rate of a chemical reaction by:

A increasing the frequency of collisions.

B decreasing the frequency of collisions.

C lowering activation energy.

D increasing activation energy.

6 An increase in the concentration will:

A increase the frequency of collisions and increase the number of successful collisions.

B increase the frequency of collisions and decrease the number of successful collisions.

C decrease the frequency of collisions and increase the number of successful collisions.

D decrease the frequency of collisions and decrease the number of successful collisions.

7 If pressure is doubled in a closed system:

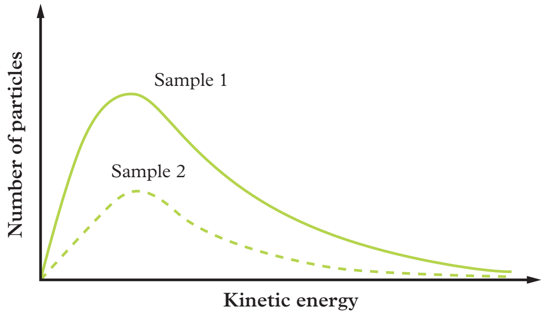
A volume is halved, causing the concentration to double and the frequency of collisions to increase.

B volume is doubled, causing the concentration to halve and the frequency of collisions to decrease.

C volume is constant, but the concentration increases, which increases the frequency of collisions.

D volume is constant, but the concentration decreases, which decreases the frequency of collisions.

8 Which of the following conclusions about the following diagram is correct?



A Sample 2 is at a higher temperature than Sample 1 because there are fewer particles that have the average kinetic energy.

B Sample 2 is at a lower concentration than Sample 1 because it has fewer particles but the same average kinetic energy.

C Sample 1 is at a higher temperature than Sample 2 because there are more particles that have the average kinetic energy.

D Sample 1 is at a lower concentration than Sample 2 because it has a greater number of particles but the same average kinetic energy.

9 Which of the following would not increase the rate of the following chemical reaction, where 50.00 mL of 0.50 M HCl is reacted with a 2.0 g piece of Mg metal ribbon at 18°C?

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

A Use a 25°C solution of HCl

B Use 2.0 M HCl

C Use 100.00 mL of 0.50 M HCl

D Use 2.0 g of powdered Mg ribbon

10 Iron reacts with oxygen in the air to form iron oxide, commonly known as rust. Cars are painted to avoid rusting, because paint protects the iron from being exposed to water and oxygen in the atmosphere. The rusting process is summarised as:

$$4Fe(s) + 3O_2(g) + 6H_2O(l) \rightarrow 4Fe(OH)_3(aq)$$

Iron(III) hydroxide will then dehydrate to form solid rust with the formula Fe_2O_3.

The process of rusting is accelerated in:

A winter environments with heavy snowfall.

B equatorial environments with higher temperatures and increased humidity.

C hot and dry desert environments.

D dry, warm bushland.

Short answer

Describe and explain

11 Define the rate of a chemical reaction. In your definition, explain how it can be measured.

12 Describe what should be on the *y*-axis and *x*-axis of a graph that represents the rate of a chemical reaction.

13 Explain the difference between the activation energy of a catalysed and an uncatalysed reaction pathway.

14 Explain the effect of increasing the temperature in an exothermic system on the rate of a chemical reaction.

15 Explain the effect of decreasing the surface area of a reactant on the rate of a chemical reaction.

16 Explain the effect of increasing the concentration of a reactant on the rate of a chemical reaction.

17 Explain how adding a catalyst affects the rate of a chemical reaction.

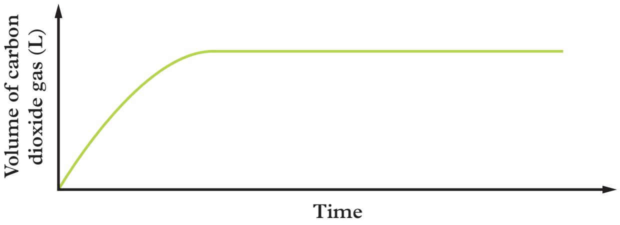
18 Draw an energy profile diagram for an exothermic reaction. On the diagram, label the enthalpy, the activation energy of an uncatalysed pathway, and the activation energy of a catalysed pathway.

19 Draw an energy profile diagram for an endothermic reaction. On the diagram, label the enthalpy, the activation energy of an uncatalysed pathway, and the activation energy of a catalysed pathway.

Apply, analyse and compare

20 A chemistry student made the claim: ‘All collisions result in a chemical reaction. When the number of collisions is increased, the rate of reaction increases.’ Explain why this student is wrong.

21 A chemical reaction between 20 mL of 1 M sodium carbonate and 5 mL of 2 M hydrochloric acid produced the following rate–time graph.



$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

- **a** The temperature of the reaction is increased and the reaction is repeated. Draw the resulting line on the graph.
- **b** The volume of sodium carbonate was increased by 10 mL in the third repeat. Draw the resulting line on the graph.
- **c** The temperature of the reaction flask is decreased. Draw the resulting line on the graph.

22 Explain how you would measure the rate of reaction for a combustion reaction. Draw a graph of the trend you would expect to observe. Identify whether this is qualitative or quantitative data.

23 Explain how you would measure the rate of reaction for a precipitation reaction. Draw a graph of the trend you would expect to observe. Identify whether this is qualitative or quantitative data.

24 Sodium sulfide is reacted with hydrochloric acid to produce sodium chloride and hydrogen sulfide gas. The gas produced smells like rotten eggs, so it is essential that it is captured rather than released into the laboratory.

$Na_2S(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2S(g)$

- **a** Explain how the gas produced by the reaction could be measured without being released into the laboratory.
- **b** Explain what you would need to use to ensure that you are measuring the rate of the gas produced.

25 Draw a Maxwell–Boltzmann diagram representing the same reaction taking place at 80°C and at 120°C.

26 1.5 g of zinc pieces is reacted with 20.0 mL of 1.0 M hydrochloric acid to form zinc(II) chloride and hydrogen gas at SLC.

- **a** Write a balanced chemical equation for this reaction.
- **b** Calculate the mass of hydrogen gas that is generated in the reaction.
- **c** Calculate the volume of the gas produced.
- **d** Plot a general graph that represents the volume of gas generated over time.
- **e** On the same graph, plot the effect of using powdered zinc in the same reaction.

27 0.55 g of calcium carbonate is reacted with 50.0 mL of 2.0 M HCl to form calcium chloride, carbon dioxide and water.

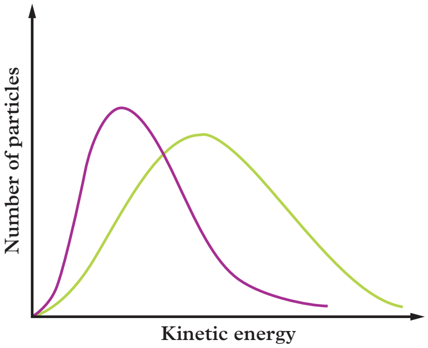
- **a** Write a balanced chemical equation for this reaction.
- **b** Calculate the mass of carbon dioxide gas that is generated in the reaction.
- **c** Calculate the volume of the gas produced.
- **d** Plot a general graph that represents the volume of gas generated over time.
- **e** On the same graph, plot the effect of an increase in HCl concentration in the same reaction.

28 Excess cadmium is reacted with 2.62 g of sulfur and 100 mL of 1.5 M hydrochloric acid according to the following equation:

$Cd(s) + S(s) + 2HCl(aq) \rightarrow CdCl_2(aq) + H_2S(g)$

- **a** Calculate the mass of hydrogen sulfide gas that is generated in the reaction.
- **b** Calculate the volume of the gas produced.
- **c** Plot a general graph that represents the volume of gas generated over time.
- **d** On the same graph, plot the effect of an increase in the temperature of the acid in the same reaction.

29 Consider the following reaction on a Maxwell–Boltzmann distribution.



a Identify what the peak of the curve represents.

b Explain why the green curve is broader and lower than the purple curve.

c Identify the curve that represents a lower temperature system.

d Copy the curve and draw a line representing the activation energy.

e Draw a second line representing the activation energy of a catalyst.

f Use the curve to explain the effect of increasing temperature on the rate of reaction.

g Use the curve to explain the effect of adding a catalyst on the rate of reaction.

Design and discuss

30 Two experiments were conducted to investigate the factors affecting the rate of a chemical reaction between magnesium metal and hydrochloric acid.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

In the first experiment, a 2 cm strip of magnesium ribbon, with a mass of 0.9 g, was reacted with 3 mL of 1 M HCl. In the second experiment, a cubic piece of magnesium, measuring 0.5 cm × 0.5 cm × 0.5 cm, with a mass of 0.9 g, was placed in 3 mL of 2 M HCl.

a Explain how the rate of reaction could be measured in these experiments.

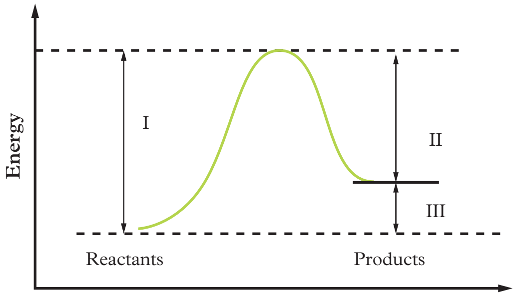
b Identify the factors being investigated in this reaction.

c Describe the major flaw associated with this experiment.

d Explain whether it is possible to determine which reaction will occur faster.

e When both reactions had finished, and no more bubbles were generated, more magnesium was added to each flask to make the reactions start again. Predict the effect of adding this extra magnesium.

31 Consider the following diagram.



a Explain whether the reaction is endothermic or exothermic.

b Identify which of the three quantities (I, II or III) is affected by the addition of a catalyst.

c Identify which of the three quantities is affected when the reaction is switched between endothermic/exothermic and endothermic/exothermic.

d Explain what would happen if the reaction was reversed. Draw an energy profile diagram to demonstrate the effect.

32 A student designs an experiment to determine the effect of temperature on the rate of reaction where a magnesium ribbon is reacted with 50 mL of 2.0 M HCl. During the experiment, the student could not find the magnesium ribbon reel and had to use scrap pieces that were unused from other experiments. The following reaction conditions were used:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

	Trial 1	Trial 2	Trial 3
Temperature (°C)	18	25	30
Mass of Mg (g)	1.30	1.15	1.05
Pieces of Mg ribbon	3	2	1
Mass of H_2 lost (g)	0.106	0.094	0.086

Critically review the student's experimental design.

a Identify the independent and dependent variables.

b State the hypothesis of the experiment.

c Identify the variables that must be controlled to effectively test the independent and dependent variables.

d Explain three improvements that must be made to the experimental design to effectively test the hypothesis.

e Discuss the accuracy of the data.

f Comment on the precision of the data.

g Discuss the validity of the experiment based on the data obtained.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER

6

Extent of chemical reactions

KEY KNOWLEDGE

- → the distinction between reversible and irreversible reactions, and between rate and extent of a reaction
- → the dynamic nature of homogeneous equilibria involving aqueous solutions or gases and their representation by balanced chemical or thermochemical equations (including states) and by concentration–time graphs
- → the change in position of equilibrium that can occur when changes in temperature or species or volume (concentration or pressure) are applied to a system at equilibrium, and the representation of these changes using concentration–time graphs
- → the application of Le Chatelier's principle to identify factors that favour the yield of a chemical reaction
- → calculations involving equilibrium expressions (including units) for a closed homogeneous equilibrium system and the dependence of the equilibrium constant (K) value on the system temperature and the equation used to represent the reaction
- → the reaction quotient (Q) as a quantitative measure of the extent of a chemical reaction: the relative amounts of products and reactants present during a reaction at a given point in time
- → responses to the conflict between optimal rate and temperature considerations in producing equilibrium reaction products with reference to the green chemistry principles of catalysis and designing for energy efficiency

Source: *VCE Chemistry Study Design (2023–2027)* reproduced by permission © VCAA

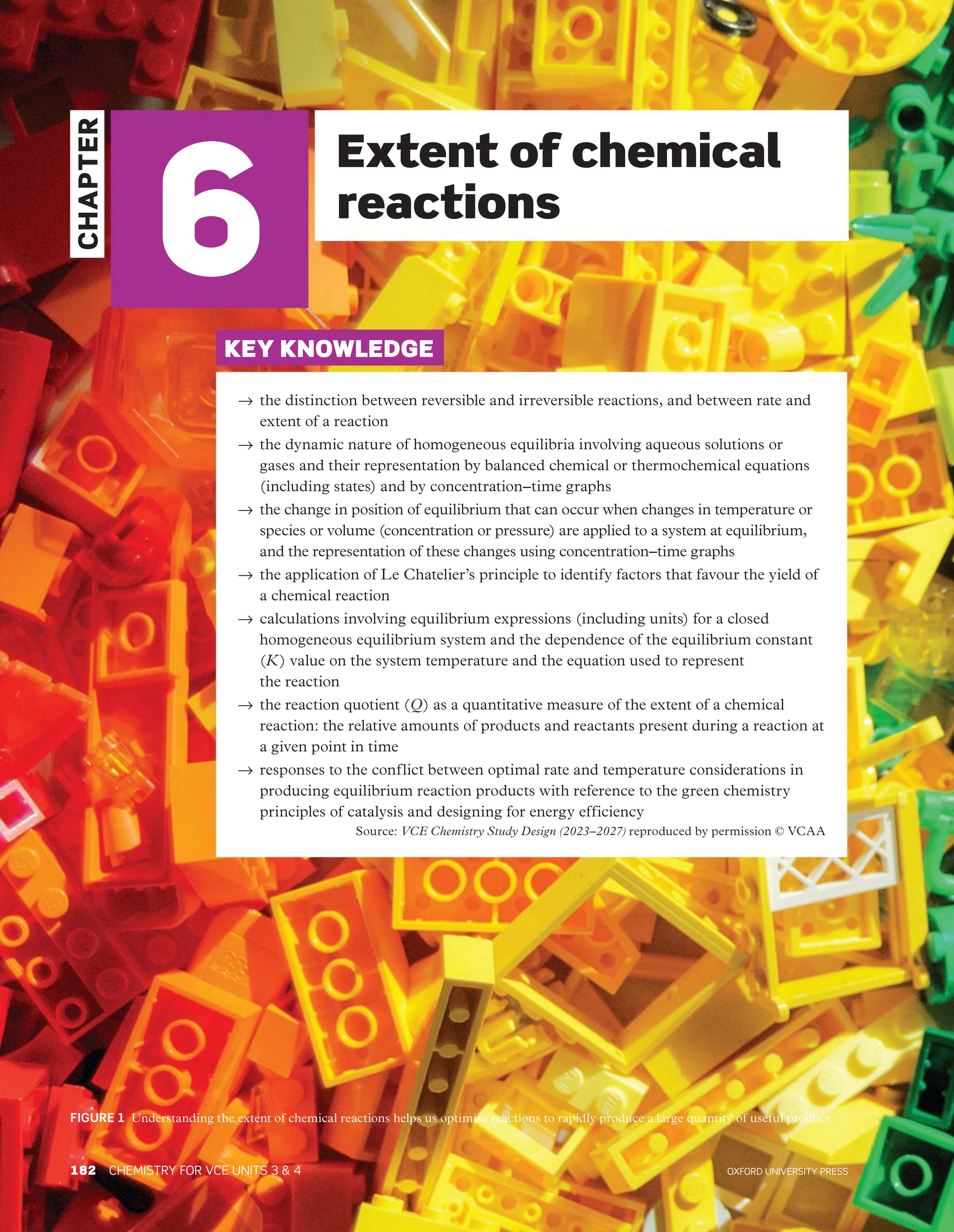


FIGURE 1 Understanding the extent of chemical reactions helps us optimise reactions to rapidly produce a large quantity of useful product.

GROUNDWORK

In Chapter 6, you will learn about reaction equilibria, the factors that influence the equilibrium, and how the equilibrium affects the extent of reactions.

This chapter will build on concepts you have already learnt in Chapter 5. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

6A Describe two factors that influence the rate of a reaction.

6A Groundwork resource
Reaction rates

6B Explain what happens to the concentration of reactants and products in a reaction.

6B Groundwork resource
Concentrations in a reaction

6C Draw an energy profile diagram for an exothermic reaction and label the energy of the reactants and products, activation energy and enthalpy.

6C Groundwork resource
Energy profile diagrams

PRACTICALS

6.3A **PRACTICAL: CONTROLLED EXPERIMENT**	What is the effect of temperature on equilibrium?	Page 508
6.3B **PRACTICAL: CONTROLLED EXPERIMENT**	What is the effect of volume, pressure and concentration on equilibrium?	pro
6.3C **PRACTICAL: CASE STUDY**	How can we stop ocean acidification?	Page 510
6.7A **PRACTICAL: SIMULATION**	What factors affect the Haber process?	pro
6.7B **PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT**	How can the formation of sulfuric acid be optimised?	pro

6.1

The rate and extent of reactions

KEY IDEAS

In this topic, you will learn that:

- irreversible reactions only proceed in the forward direction (reactant → product), whereas reversible reactions can proceed in both forward and reverse directions (reactant ⇌ product)
- the rate of reaction is how quickly the forward reaction occurs, whereas the extent of a reaction is the degree to which the forward reaction proceeds.

Before a chemical reaction occurs, there are only reactants. As the reaction occurs, the amount of reactant decreases and the amount of product increases.

In Chapter 5, you learnt about the rate of reactions and the factors that influence the reaction rate. Now, you will explore the reversibility of chemical reactions (the ability for products to be converted back into reactants) and reaction extent (the degree to which a reaction proceeds).

Irreversible and reversible reactions

The products of some chemical reactions escape into the atmosphere. For example, a combustion reaction in an **open system** (where matter can be lost to the atmosphere) produces gaseous products. When ethanol reacts with oxygen, it combusts in a highly exothermic reaction to produce carbon dioxide and water vapour according to the following balanced **forward reaction**:

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

The ethanol and oxygen are consumed in the reaction, and the gases that form are lost to the atmosphere. This reaction will continue in the forward direction to produce more products until there are no more reactants. Since the products cannot be turned back into their starting reactants, this is called an **irreversible reaction**. Irreversible reactions are represented using unidirectional arrows (→).

Some reactions can proceed both forwards and backwards. At any one time, there is a mixture of reactants and products in the system. This means that it is a **reversible reaction**. Reversible reactions are represented using bidirectional arrows (⇌) to show the forward and **reverse reactions** (Figure 2). You may recognise these from the electrochemical series in Chapter 4.

A significant number of reactions involve keeping reactants and products contained in the same reaction vessel. These are called **closed systems**. They prevent products from escaping once they are formed, since matter is contained within the system. Reversible reactions can occur in both open and closed systems.

FIGURE 1 Ethanol reacts with oxygen in an irreversible combustion reaction.

open system
a reaction vessel that has no lid, meaning that reactants or products can be lost to the atmosphere

forward reaction
the reaction between reactants to form products

irreversible reaction
a reaction that can only proceed in the forward direction

reversible reaction
a reaction that can proceed in both forward and reverse directions

reverse reaction
the backward reaction where the products react to re-form the reactants

closed system
a reaction vessel that is closed off, or has a lid, meaning that no reactant or product particles can escape

A great way to grasp the concept of reversible reactions is by looking at thermochemical reactions. Figure 3a shows the energy profile for an exothermic reaction. The activation energy for the reaction is 40 kJ. The change in enthalpy is the difference between the chemical energy of the products (20 kJ) and the chemical energy of the reactants (40 kJ): –20 kJ.

If this reaction is reversed, the products of the forward reaction would now be the reactants of the reverse reaction. Therefore, the energy profile will be reversed (Figure 3b). The reverse reaction has an activation energy of 60 kJ and a change in enthalpy of 20 kJ.

If the chemical species in the system can collide with enough energy to overcome the activation energy for both the forward and reverse reactions, the reaction can proceed forwards and backwards.

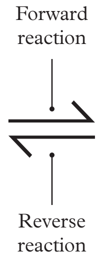


FIGURE 2 The reversible reaction arrow

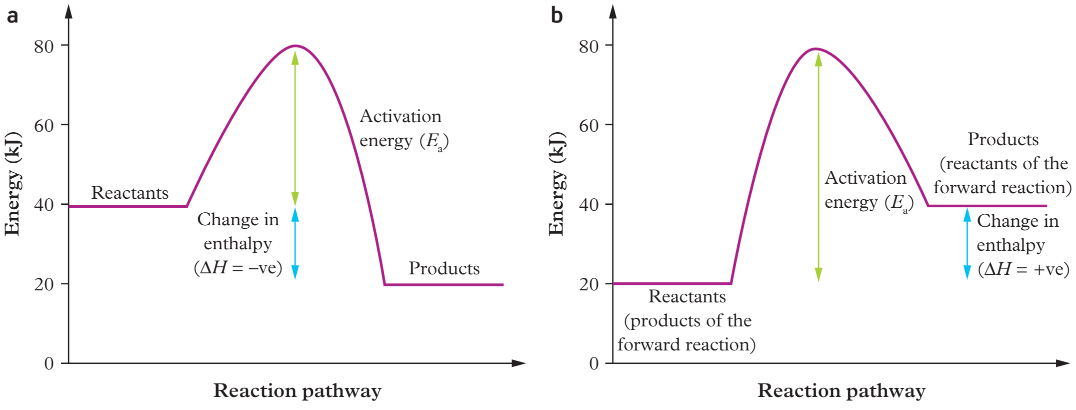


FIGURE 3 Energy profile diagrams for **a** an exothermic reaction and **b** the reverse endothermic reaction

Study tip

The electrochemical series shows reversible redox half-equations. The reduction half-reactions are the forward reactions and the oxidation half-reactions are the reverse reactions.

Rate and extent of a reaction

In chemical industries, such as mining, the goal is to increase the amount of product (or yield) of a reaction. This allows a chemist to obtain more of the useful product, which can be sold or used to perform another reaction. Much of the innovation and design of new processes or optimisation of existing ones is to try and get as *much* product as possible, as *fast* as possible.

As you learnt in Chapter 5, the speed at which a product is formed is called the rate of a reaction. This is the change in the concentration of reactants or products per unit of time. The amount of product formed in a specific amount of time can be increased by an increase in concentration, surface area, pressure or temperature, a decrease in volume, and by the addition of a catalyst. In closed systems containing reversible reactions, the rate of reaction can also be increased; however, this will increase the rate of both forward and reverse reactions.

The **extent of a reaction** refers to the magnitude of the reaction. The yield is a measure of this extent. Specifically, it is the amount of product that can be obtained from a reaction. To increase the extent of the reaction means to increase the yield of product that can be obtained from it.

extent of a reaction
a measure of the magnitude to which a reaction proceeds in the forward direction

FIGURE 4 The *extent* of crystal formation is how much has formed and the *rate* is how quickly it has formed.

6.1 SKILL DRILL

Analysing and evaluating the yield of an acid–base reaction

Key science skill: Analyse and evaluate data and investigation methods

The following reaction occurs in a beaker:

$$HCl(aq) + NaHCO_3(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

In the experiment, 50 mL of 0.33 M $NaHCO_3$ was reacted with 100 mL of 0.5 M HCl in a beaker. On completion of the reaction, the system was found to have lost a mass of 0.71 g. When the reaction was repeated two more times, it lost masses of 0.70 g and 0.72 g.

Practise your skills

1 Explain why the reaction lost mass.

2 Calculate the theoretical mass that can be lost based on the quantities of the reactants.

3 Comment on the accuracy of the experimental repeats – i.e., state which is more accurate. Use the data to support your answer.

4 Comment on the precision, repeatability and reproducibility of the data.

Need help analysing and evaluating data? See Topic 1.8 (page 24).

6.1 CHECK YOUR LEARNING

Describe and explain

1 Explain what is required for a reverse reaction to occur.

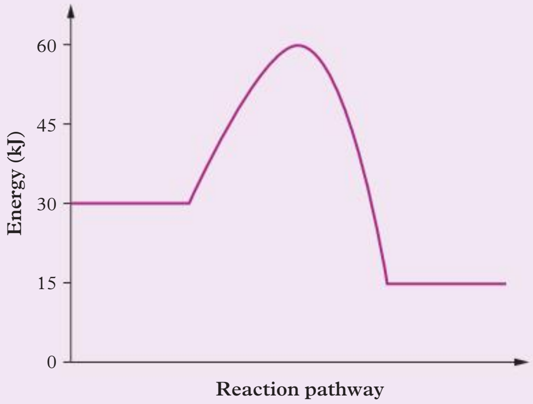
2 Describe the difference between rate and extent of a reaction.

Apply, analyse and compare

3 Compare reversible and irreversible reactions, in terms of the amounts of reactants and products in the reaction system.

4 Predict what will happen to the yield of product if the rate of reaction is increased.

5 An energy profile diagram is shown below for an exothermic reaction.



a Calculate the change in enthalpy of the reaction.

b Calculate the activation energy of the reaction.

c Draw an energy profile diagram for the reverse reaction.

d Calculate the activation energy of the reverse reaction.

e The products formed from the forward reaction have 40 kJ of energy. Determine whether this reaction is reversible and justify your response.

6.2 The nature of homogeneous equilibria

KEY IDEAS

In this topic, you will learn that:

- chemical reactions that can proceed both in both forward and reverse direction are called equilibrium reactions
- dynamic equilibrium occurs when the rate of the forward and backwards reactions are equal
- homogeneous equilibrium systems occur when all reactants and products are the same state
- equilibrium systems can be represented by concentration–time graphs.

In the last topic, you learned that reversible reactions are represented using bidirectional arrows (⇌). These are also called **equilibrium arrows**. Reversible reactions are therefore also called **equilibrium reactions**.

In equilibrium reactions, both forward and reverse reactions occur simultaneously within a closed system. This means that, at any time, both reactants and products are present in the system. The reaction does not go to completion (where only products exist).

One example of an equilibrium reaction is the partial dissociation of weak acids, such as acetic (ethanoic) acid, in water:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

equilibrium arrows bidirectional arrows that show that a reaction is reversible

equilibrium reaction a reaction that is reversible; reactants and products are both present in the system at any one time

Dynamic equilibrium

The reaction reaches a state of **dynamic equilibrium** when there is a perfect balance between forward and reverse reactions. In other words:

- the reaction is incomplete – it does not consist of only products
- bonds are being broken and re-formed constantly; particles are constantly changing between reactant and product – 'dynamic'
- the rates of the forward and reverse reactions are equal – 'equilibrium'.

dynamic equilibrium the state a reaction reaches when the forward and reverse reactions are occurring at the same rate

Study tip

Remember that for equilibrium reactions to occur, reactants must be able to form products, and products must be able to be converted back into reactants. This can only occur in a closed system where matter is not lost to the environment.

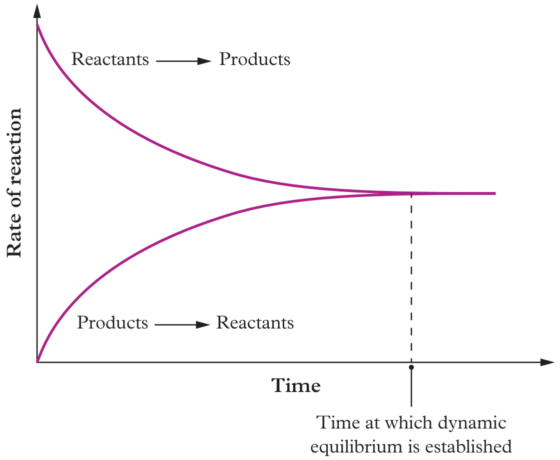


FIGURE 1 Dynamic equilibrium is reached when the rates of the forward and reverse reactions are equal.

Consider the synthesis reaction between nitrogen and hydrogen to form ammonia:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.3\,kJ\,mol^{-1}$$

6.2 Worked example
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6.2 Worked example
Video demonstration

At dynamic equilibrium, nitrogen, hydrogen and ammonia are present in the reaction vessel, although not necessarily in equal amounts or concentrations. As the bonds in the nitrogen and hydrogen molecules are breaking, they are re-forming as ammonia. At the same, as ammonia is forming, its bonds are also breaking to re-form the nitrogen and hydrogen molecules. The rate of formation of ammonia at equilibrium is equal to the rate of formation of nitrogen and hydrogen. See how to write a balanced equilibrium reaction equation in Worked example 6.2.

Homogeneous equilibria

homogeneous equilibrium
an equilibrium reaction where the states of all reactants and products are the same

heterogeneous equilibrium
an equilibrium reaction where the states of reactants and products are not the same

A **homogeneous equilibrium** system, *homo* meaning ‘same’, is one in which all of the reactants and products are present in the same state. A **heterogeneous equilibrium** system, *hetero* meaning ‘different’, is one in which the reactants and products are in different states.

For example:

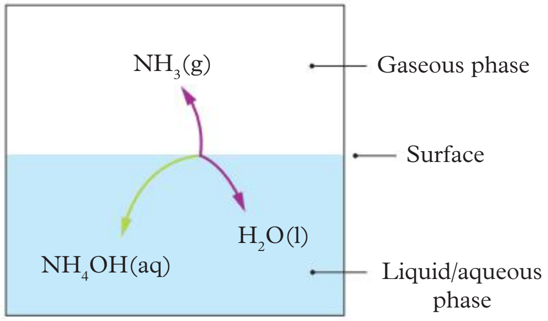
Homogeneous: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92.3\,kJ\,mol^{-1}$

Heterogeneous: $NH_4OH(aq) \rightleftharpoons NH_3(g) + H_2O(l)$ $\Delta H = -14.8\,kJ\,mol^{-1}$

In homogeneous equilibria, all reactants and products exist in either a gaseous, liquid or aqueous state. This maximises the ability of particles to collide successfully, increasing not only the rate but the extent and yield of the reaction.

Heterogeneous systems work very differently to homogeneous systems because of the difference in states.

For example, NH_4OH in the aqueous phase will dissociate into liquid H_2O and NH_3 gas (Figure 2). Once the NH_3 is formed, it moves to the gas phase and can only react in the reverse reaction if it is in contact with the water in the aqueous solution. At any time, not all NH_3 molecules in the system will be available at the surface of the water. Therefore, there is a limited number of particles available to react in the forward or reverse reactions Explore another example in Real-world chemistry 6.2.



$NH_3OH(aq) \rightleftharpoons NH_3(g) + H_2O(l)$

FIGURE 2 In a heterogeneous equilibrium system, particles in the gas phase can only interact with particles in the liquid phase if they are located at the surface.

6.2 Real-world chemistry
Find me in your obook pro

Concentration–time graphs

The change in concentrations of reactants and products as a reaction proceeds can be represented using concentration–time graphs. These also allow you to see when dynamic equilibrium is established.

Because the rates of the forward and reverse reactions are equal at dynamic equilibrium, the concentrations of the reactants and products stabilise.

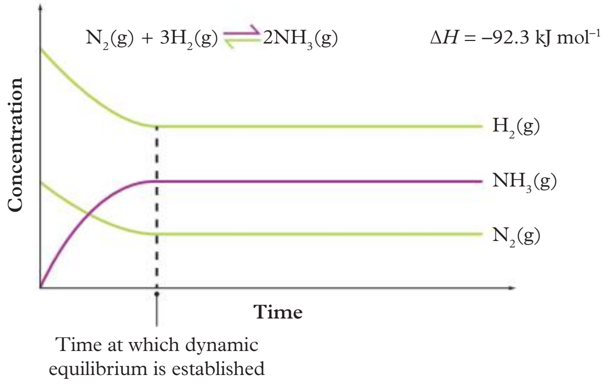


FIGURE 3 Once equilibrium is reached, concentrations of reactants and products remain constant.

6.2 CHECK YOUR LEARNING

Describe and explain

1 Describe what is required for a reaction to have reached dynamic equilibrium.

2 Explain why a homogeneous equilibrium system is favoured over a heterogeneous system.

Apply, analyse and compare

3 Determine whether each of the following reactions are a homogeneous or heterogeneous equilibrium reaction.

  a $2HF(aq) \rightleftharpoons 2H^+(aq) + 2F^-(aq)$

  b $2PbO(s) + O_2(g) \rightleftharpoons 2PbO_2(aq)$

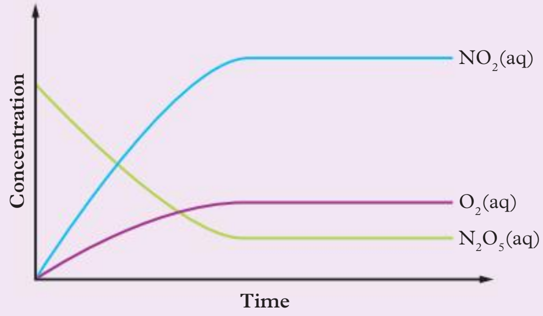
  c $CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

  d $CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$

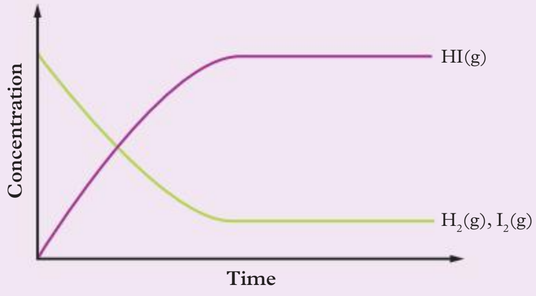
4 Nitrogen and hydrogen are placed in a sealed glass jar. After 30 seconds, the jar is knocked to the floor, smashing it. Determine what happens to the concentration of the reactants and products before and after the accident.

5 Look at the concentration–time graphs I–IV.

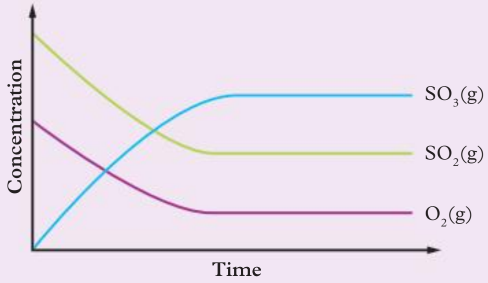
I



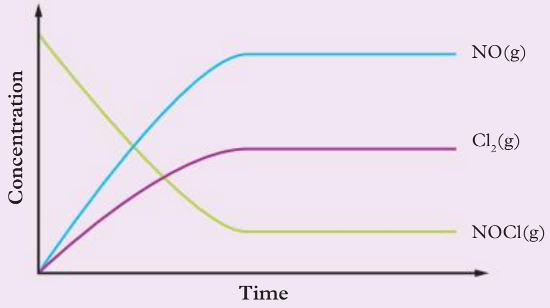
II



III



IV



For each graph:

  a Identify the reactants and products of the forward reaction.

  b Write a balanced equilibrium equation for the reaction.

Design and discuss

6 A student in your class claims that Earth is an open equilibrium system. Evaluate this claim and communicate whether you believe it to be correct. Justify your argument.

6.3 Changes in the position of equilibrium

KEY IDEAS

In this topic, you will learn that:

- the position of equilibrium is the relative amounts of reactants and products at equilibrium
- the position of equilibrium can be changed by adding or removing a species, changing temperature, changing pressure or volume in gaseous systems, or diluting aqueous solutions
- Le Chatelier's principle states that if the position of equilibrium in a system is subjected to change, the system will shift to partially oppose the effect of the change.

Chemists must have a good understanding of the factors affecting a chemical reaction and how these factors can be manipulated to maximise the amount of desirable product from a reaction. The aim is to maximise the yield of a chemical reaction without losing or destroying any materials in the process.

Before we look into methods to improve reaction yield, we need to understand how these factors influence the behaviour of chemicals in a reaction system. This topic focuses on the work of French chemist and engineer Le Chatelier, who is best known for his work on equilibrium.

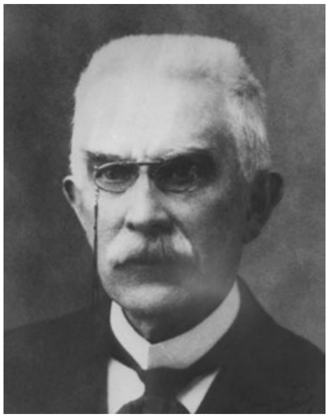


FIGURE 1 French chemist Henry Louis Le Chatelier (1850–1936) developed Le Chatelier's principle, which predicts the effect of changes on a chemical equilibrium.

Le Chatelier's principle

In an equilibrium system, forward and reverse reactions occur simultaneously. An equilibrium reaction can be manipulated to be pushed forward, increasing the concentration of the products in a system, or backward, increasing the concentration of reactants.

The factors that can change the concentrations of reactants and products are:

- the addition or removal of a species (reactants or products)
- a change in temperature
- a change in pressure and volume in gaseous systems
- the dilution of aqueous solutions.

By manipulating these factors, we change the **position of equilibrium**, meaning we change whether the forward or reverse reaction is favoured at equilibrium. Put simply, we can change the relative amounts of reactants and products at equilibrium.

Le Chatelier developed a principle to predict the effect of these changes on a system at equilibrium. **Le Chatelier's principle** states that if the position of equilibrium in a system is subjected to change, the system will shift to partially oppose the effect of the change and re-establish equilibrium.

position of equilibrium the relative amounts of reactants and products in a reaction system at equilibrium

Le Chatelier's principle if the position of equilibrium in a system is subjected to change, the system will shift to partially oppose the effect of the change

6.3A

PRACTICAL: CONTROLLED EXPERIMENT

What is the effect of temperature on equilibrium? Go to page **508**.

6.3B

PRACTICAL: CONTROLLED EXPERIMENT

What is the effect of volume, pressure and concentration on equilibrium? Go to your **Student obook pro**.

The effect of adding or removing a species

When the concentration of a reactant or product is increased, the system partially opposes this change by moving to decrease its concentration. If the concentration of a reactant or product is decreased, the system partially opposes the change by moving to increase its concentration. This is summarised in Table 1.

Study tip

You must use the terms *shift*, *partially oppose* and *net forward/reverse reaction* to earn full marks in your exam.

TABLE 1 Effects of adding or removing species on the position of equilibrium

What has happened?	How does the system partially oppose the change?	What is the effect on the position of equilibrium?
Reactant is added.	It shifts to decrease the amount of reactant.	Net forward reaction →
Product is added.	It shifts to decrease the amount of product.	Net reverse reaction ←
Reactant is removed.	It shifts to increase the amount of reactant.	Net reverse reaction ←
Product is removed.	It shifts to increase the amount of product.	Net forward reaction →

The effect of adding or removing reactants or products from a system at equilibrium can be shown using concentration–time graphs. You can see this in Figure 2, which is an equilibrium system with the following equation:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Each species begins at equilibrium (black horizontal lines). When a species is added or removed, a vertical line is drawn. In this case, H_2 is added. Because H_2 is a reactant, the system partially opposes the change to decrease the amount of reactant. This means that there is a **net forward reaction**. The concentrations of reactants N_2 and H_2 decreases and the concentration of the product NH_3 increases (green curves) until equilibrium is re-established (blue horizontal lines). You can tell that there has been a partial change because the final concentrations, after equilibrium is re-established, are not the same as the initial equilibrium concentrations.

We look at the mole ratios from the balanced equation to determine the relative amounts of change. N_2 changes by a factor of 1, H_2 changes by a factor of 3 and NH_3 changes by a factor of 2 (green arrows).

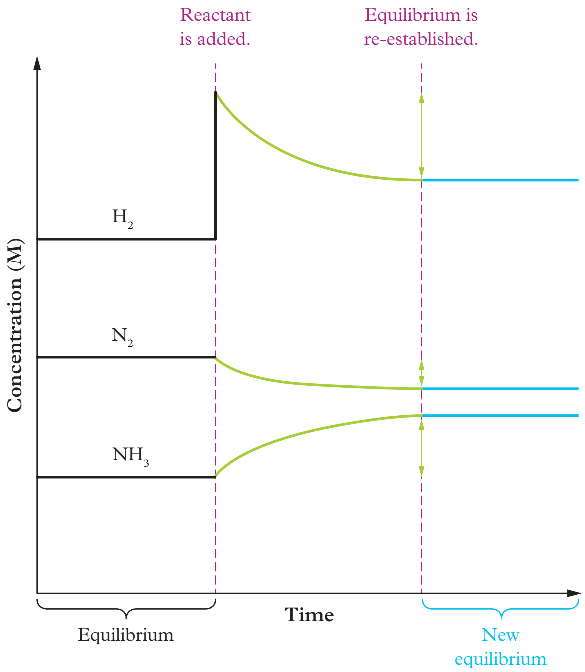


FIGURE 2 The effect of adding a reactant to an equilibrium system

The effect of changing the temperature

Temperature is a measure of heat energy. When we increase the temperature of a system, we also increase the energy of the species in the reaction. All reactions require energy to overcome their activation energy. However, the amount of energy depends on the size of the activation energy and whether the reaction is exothermic or endothermic.

net forward reaction
a shift in the position of equilibrium in which more products are being formed compared to reactants being re-formed

6.3C
PRACTICAL: CASE STUDY

How can we stop ocean acidification?
Go to page **510**.

Endothermic reactions

In an endothermic reaction, energy must be absorbed from the environment for the reaction to proceed (Figure 3).

To determine the effect of temperature on a reaction, let's consider the energy to be a reactant or product. In an endothermic reaction, energy can be considered as a reactant:

$$\text{energy} + \text{reactants} \rightleftharpoons \text{products}$$

When we have increased the temperature, we can treat this as though we've *added* a reactant: energy. This will favour the formation of products. Le Chatelier's principle states that the system will oppose the effect of this change, so the reaction will shift to *consume* the extra energy. This will favour the forward reaction or cause a net forward reaction to re-establish equilibrium.

What about if we cooled the system down? We would be *removing* energy as a reactant. The system would shift to *produce* more energy. This will favour the reverse reaction or cause a **net reverse reaction** to re-establish equilibrium.

net reverse reaction
a shift in the position of equilibrium in which more reactants are being re-formed compared to products being formed

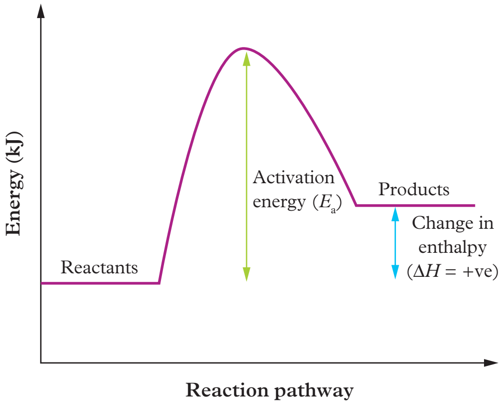


FIGURE 3 An endothermic reaction energy profile

Exothermic reactions

In an exothermic reaction, reactants have more energy than products. This means that energy is released into the environment and the change in enthalpy is negative. In an exothermic reaction, energy is a product:

$$\text{reactants} \rightleftharpoons \text{products} + \text{energy}$$

Think about how an increase or decrease in temperature will affect the position of equilibrium in an exothermic reaction. The effects are summarised in Table 2, along with responses to temperature change in endothermic reaction systems.

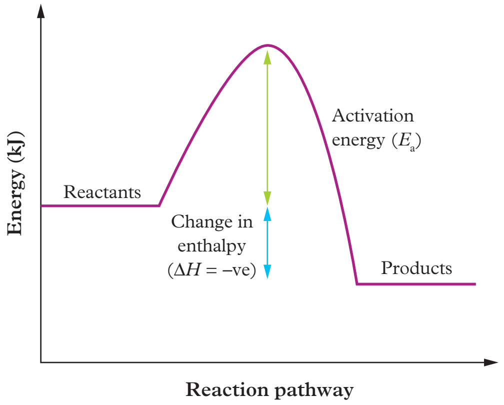


FIGURE 4 An exothermic reaction energy profile

TABLE 2 Effects of temperature change on the position of equilibrium

The reaction system is...	What has happened?	How does the system partially oppose the change?	What is the effect on the position of equilibrium?
endothermic	Temperature in an endothermic system has increased (energy is added).	It shifts to consume the extra energy.	Net forward reaction →
	Temperature in an endothermic system has decreased (energy is removed).	It shifts to produce more energy.	Net reverse reaction ←
exothermic	Temperature in an exothermic system has increased (energy is added).	It shifts to consume the extra energy.	Net reverse reaction ←
	Temperature in an exothermic system has decreased (energy is removed).	It shifts to produce more energy.	Net forward reaction →

The effect of changes in temperature in an equilibrium system on the concentration of the reactants and products can also be shown using concentration–time graphs. For example, consider the following reaction:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) \qquad \Delta H = +180.5\,\text{kJ}\,\text{mol}^{-1}$$

As before, each species begins at equilibrium. But unlike Figure 2, there is no vertical black line in the curve (Figure 5). In fact, temperature is the only change where there is no initial vertical change in the graph. To determine what happens to the concentrations when the temperature is increased, we have to first check whether the reaction is endothermic or exothermic. Since the change in enthalpy is positive, this is an endothermic reaction.

When temperature is increased, energy is added to the system. In an endothermic reaction, energy is a reactant. The system will partially oppose this change by trying to reduce energy and there will be a net forward reaction to help re-establish equilibrium.

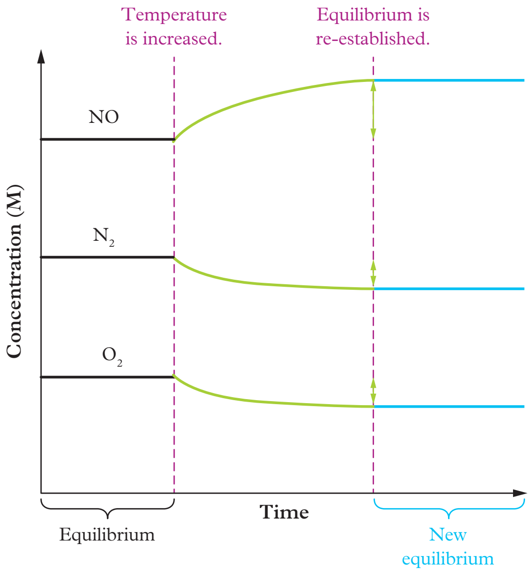


FIGURE 5 The effect of increasing temperature of an equilibrium system

This means that the concentration of NO will increase, while the concentrations of N_2 and O_2 will decrease. Look at the mole ratios to determine the relative amounts of change. N_2 and O_2 both change by a factor of 1, and NO changes by a factor of 2.

Worked example 6.3A shows you how to determine what happens to the position of equilibrium in a thermochemical reaction.

6.3A WORKED EXAMPLE

DETERMINING THE EFFECT OF DECREASING TEMPERATURE ON EQUILIBRIUM

Phosphorus trichloride and chlorine gas are in equilibrium with phosphorus pentachloride:

$$PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g) \qquad \Delta H = -92.5\,\text{kJ}\,\text{mol}^{-1}$$

Determine the effect that a decrease in temperature of the system will have on the position of equilibrium.

Solution

Think	Do
Step 1: Identify whether the forward reaction is endothermic or exothermic and use this to determine whether energy is a reactant or product.	The change in enthalpy is negative, so this is an exothermic reaction. Energy is a product.
Step 2: Use Le Chatelier's principle to determine what happens to the position of equilibrium when temperature is decreased.	A decrease in temperature means a decrease in energy as a product. To partially oppose this change, the position of equilibrium will shift to increase energy (a product), so there will be a net forward reaction.

Study tip

Treat energy as a reactant in endothermic reactions (positive ΔH) and as a product in exothermic reactions (negative ΔH).

The effect of changing pressure and volume in gaseous systems

pressure
the force exerted, per unit area, by one substance on another substance

Study tip

Doubling the volume will exactly halve the pressure. Halving the volume will exactly double the pressure.

Gases can exist in fixed-volume and variable-volume systems. According to Charles' law, as the volume of a system decreases, the **pressure** increases. More collisions between the particles and the walls of the container, as they are contained in a smaller space, cause the concentrations of reactant and product particles to increase.

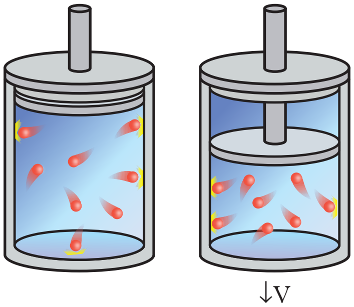


FIGURE 6 When volume is decreased, pressure increases.

Predicting the effect of changing pressure and volume in gaseous systems is not as straightforward as predicting the effect of adding or removing species or changing the temperature. Equilibrium will shift depending on the number of collisions or the amount of particles.

Let's look at the following reaction as an example:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

When the volume is increased, the pressure decreases. There are fewer collisions as the space between particles is increased (so their concentration decreases). Instead of decreasing the volume again, the system will shift to oppose this change by increasing the number of particles.

To determine the change in the position of equilibrium, we look at the total number of particles on each side of the reaction (the coefficients in the balanced reaction). On the reactant side, there is 1 molecule (or particle) of N_2O_4. On the product side, there are 2 molecules (or particles) of NO_2. There are more particles on the product side, so there will be a net forward reaction (Figure 7).

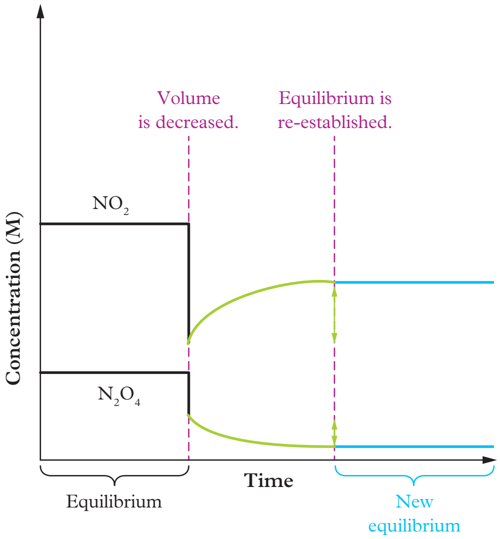


FIGURE 7 The effect of changing pressure and volume of an equilibrium system

TABLE 3 Effects of changing volume and pressure of a gaseous system on the position of equilibrium

What has happened?	How does the system partially oppose the change?	What is the effect on the position of equilibrium?
The volume decreases and pressure increases.	It shifts to the side with fewer particles.	It depends on the reaction equation.
The volume increases and pressure decreases.	It shifts to the side with more particles.	It depends on the reaction equation.

In a fixed-volume system, volume cannot change. Any changes to pressure in this type of system are not the result of changing the volume. Instead, pressure in fixed-volume systems can be increased by adding a reactant or product, or increasing temperature.

The addition of an inert (unreactive) gas, such as argon, can also increase the pressure of the system. However, this will have no effect on equilibrium because the inert gas does not interact with the reactants and products, so it will not affect the number of collisions. Therefore, the addition of an inert gas to a gaseous equilibrium system has no effect.

6.3B WORKED EXAMPLE

DETERMINING THE EFFECT OF DECREASING VOLUME ON EQUILIBRIUM

A gaseous system containing sulfur dioxide, oxygen and sulfur trioxide reacts according to the following equation:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

Determine the effect that halving the volume of the system will have on the position of equilibrium.

Solution

Think	Do
Step 1: Use Charles' law to determine the effect of halving the volume on the number of collisions or concentration of particles.	Halving (decreasing) the volume will increase the pressure. There will be more collisions and the concentration of particles will increase.
Step 2: Use Le Chatelier's principle to determine what happens to the position of equilibrium when the number of collisions and particles is increased.	To partially oppose this change, the position of equilibrium will shift to reduce the number of particles.
Step 3: Look at the coefficients in the balanced reaction to see which side has fewer particles.	On the reactant side, there are two SO_2 particles and one O_2 particle (three particles in total). On the product side, there are two SO_3 particles. The product side has fewer particles, so there will be a net forward reaction.

The effect of dilution in aqueous systems

Aqueous systems can be diluted simply by adding water. This decreases the concentration of all reactants and products.

Similar to when we change volume and pressure, the position of equilibrium will shift depending on the number of particles on each side of the reaction. When an aqueous reaction system is diluted, the number of collisions between particles decreases. Le Chatelier's principle says that the system will shift to partially oppose this change by moving to increase the number of particles. Therefore, equilibrium will shift in the direction that forms more particles.

6.3 CHECK YOUR LEARNING

Describe and explain

1 Describe Le Chatelier's principle.

2 Explain how volume and pressure affect the concentration of particles in a fixed-volume reaction system.

Apply, analyse and compare

3 Compare the effect of the position of equilibrium when the temperature is increased in an endothermic reaction system to when temperature is increased in an exothermic reaction system.

4 Explain the effect of the following changes on the equilibrium that exists between sulfur dioxide, oxygen and sulfur trioxide:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \quad \Delta H = -197\,kJ\,mol^{-1}$$

a increasing the volume

b decreasing the temperature

c increasing the amount of sulfur trioxide

d increasing the temperature.

Design and discuss

5 Discuss why the word *net* is used to describe the direction of change in an equilibrium system.

6 Discuss why the expression *partially oppose* is used when describing the effect of a change in equilibrium.

6.4 Using Le Chatelier's principle to increase yield

KEY IDEAS

In this topic, you will learn that:

- yield is the amount of product obtained from a reaction
- Le Chatelier's principle can be applied to maximise the yield of a chemical reaction.

yield
the amount of product obtained from a chemical reaction

Chemical **yield** is the amount of desired product that is generated by a chemical reaction. Le Chatelier laid the foundation for industrial chemistry because his theory was able to explain how to optimise chemical yield.

FIGURE 1 Weighing the products of a reaction to determine the yield

Yield of a chemical reaction

You might remember from Topic 6.1 that the yield is a measure of the extent of a reaction (how much product is formed). We can use Le Chatelier's principle and our understanding of the factors that affect the position of equilibrium to favour the forward reaction. In other words, we can change the conditions in which an equilibrium reaction occurs to produce a higher quantity of product.

Table 1 below summarises how you could manipulate an equilibrium system to cause a net forward reaction.

TABLE 1 Changes to an equilibrium system that favour the yield of the chemical reaction

Change	Types of systems	How does the system partially oppose the change?
Addition of reactants	All reaction systems	It shifts to decrease the amount of reactant, therefore increasing the amount of product.
Removal of products	All reaction systems	It shifts to increase the amount of product.
Increase in temperature	Endothermic reactions	It shifts to consume the extra energy (reactant), therefore increasing the amount of product.
Decrease in temperature	Exothermic reactions	It shifts to produce more energy (product), therefore increasing the amount of product.
Decrease in volume and increase in pressure	Gaseous systems with more particles on the reactant side	It shifts to the side with fewer particles (fewer collisions), therefore increasing the amount of product.
Increase in volume and decrease in pressure	Gaseous systems with more particles on the product side	It shifts to the side with more particles (more collisions), therefore increasing the amount of product.
Dilution by adding water	Aqueous systems	It shifts to the side with more particles, therefore increasing the amount of product.

theoretical yield
the maximum amount of product formed in a chemical reaction

experimental yield
the actual amount of product formed in a chemical reaction; also called actual yield

Theoretical and actual yield

The **theoretical yield** is the maximum predicted quantity of a product that can be obtained from a reaction. The actual yield of a reaction is usually less than this theoretical value. This is called the **experimental yield**. It is the quantity of product that is actually obtained from a reaction.

The efficiency of a chemical reaction (conversion of reactants to products) can be expressed by the **percentage yield**:

$$\text{percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100\%$$

Worked example 6.4 shows you how to use your understanding of Le Chatelier's principle to increase the yield of a reaction.

percentage yield
the degree to which reactants are converted to products

6.4 WORKED EXAMPLE

MAXIMISING THE YIELD OF CHEMICAL REACTIONS

Consider the following equation:

$$2NO_2(g) \rightleftharpoons N_2O_4(g) \qquad \Delta H = -25\,\text{kJ}\,\text{mol}^{-1}$$

State all of the factors which increase the yield of the reaction and explain how they do so.

Solution

Think	Do
Step 1: Consider the addition of a reactant or product.	If the concentration of the NO_2 reactant is increased, the system will partially oppose the change by moving net forward, increasing the yield of N_2O_4.
Step 2: Consider changing the pressure.	If the pressure is increased by decreasing the volume, the concentration of both NO_2 and N_2O_4 is increased. This increase in concentration is partially opposed by moving to the side with less particles, resulting in a net forward reaction and increasing the yield of N_2O_4.
Step 3: Consider changing the temperature.	If the temperature is decreased in an exothermic system, the system partially opposes the change by moving net forward to release energy from reactant bonds. This will increase the yield of N_2O_4.
Step 4: Consider dilution.	This reaction is gaseous and not aqueous, so it cannot be diluted.
Step 5: Consider adding a catalyst.	A catalyst will increase the rate of the forward and reverse reactions and make the reaction reach equilibrium faster, but will not affect the yield of the reaction.

6.4 CHECK YOUR LEARNING

Describe and explain

1 Explain the meaning of the term *yield*.

2 When you are asked to discuss the effect of a change on the yield of a reaction, identify which net direction you should explain. Justify your response.

3 Explain the impact on equilibrium yield if the desired product was continually removed throughout the reaction.

Apply, analyse and compare

4 Discuss the factors that can increase the yield of the following reactions. You must include a justification for your answer.

- **a** $2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g)$
- **b** $2N_2O_5(aq) \rightleftharpoons 4NO_2(aq) + O_2(aq)$
- **c** $H_3PO_4(aq) \rightleftharpoons PO_4^{3-}(aq) + 3H^+(aq)$
- **d** $4NO(g) + 6H_2O(l) \rightleftharpoons NH_3(g) + 5O_2(g)$

5 Discuss the factors that can increase the yield of the following reactions. You must include a justification for your answer.

- **a** $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H < 0$
- **b** $N_2(g) + 2O_2(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H > 0$
- **c** $2NO_2(g) + O_2(g) \rightleftharpoons N_2(g) + 2O_3 \qquad \Delta H > 0$

Design and discuss

6 NO_2 is a dark brown gas that exists in equilibrium with N_2O_4, a colourless and highly toxic gas. The mixture of gases is placed in a large syringe and left to reach equilibrium. A student pushed the plunger in to determine the effect of decreasing the volume. Initially the mixture becomes darker and then gradually lighter. Using Le Chatelier's principle, evaluate this effect and explain why this happens.

6.5

Equilibrium expressions

KEY IDEAS

In this topic, you will learn that:

- the equilibrium law (represented by the equilibrium constant, *K*) is used to calculate the ratio of reactants to products in an equilibrium system
- *K* is used to determine the position of equilibrium
- the only factor that can change a *K* value is an increase or decrease in temperature.

As the theory of equilibrium was developed, scientists quickly realised that there was a relationship between the concentration of reactants and products in a reaction system.

Equilibrium law

The equilibrium law is used to determine the mathematical relationship between the concentration of reactants and products, called the **equilibrium constant**, *K*. All concentrations are to the power of their coefficients. The concentrations of the products at equilibrium are multiplied together and divided by the concentrations of the reactants at equilibrium.

equilibrium law the concentration of products to the power of their coefficients, divided by the concentration of products to the power of their coefficients, is equal to the equilibrium constant, *K*

equilibrium constant (*K*) the ratio of reactants to products in a reaction when it is at equilibrium

equilibrium expression the non-numerical representation of the equilibrium law, which states the chemicals to the power of their coefficients

A chemical reaction can be summarised by the general equation:

$$wA + xB \rightleftharpoons yC + zD$$

where A, B, C and D are the reactants and products, and w, x, y and z are their coefficients.

The equilibrium law is therefore expressed as:

$$K = \frac{[C]^y[D]^z}{[A]^w[B]^x}$$

This **equilibrium expression** can only be applied to closed homogeneous equilibrium systems where all reactants and products are in the same state, and there is no loss of matter to the atmosphere.

Calculating *K*

Equilibrium is one of the few concepts in chemistry where the units change based on the chemical reaction used in the calculation. As equilibrium is represented as a concentration fraction, units are also in concentration.

For the reaction $CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O$, when calculating the *K* value, the units of concentration must be considered:

$$K = \frac{[CH_4][H_2O]}{[CO][H_2]^3}$$

If the molecules are removed so that only the units of concentration are represented in the fraction, the equation becomes:

$$K = \frac{M \times M}{M \times (M \times M \times M)} = \frac{1}{M \times M} = M^{-2}$$

As there are two M units in the numerator and 4 in the denominator, we can cancel out the numerator, leaving two in the denominator. This makes the unit of the *K* value M^{-2}.

Study tip

K is calculated using the concentrations of each of the reactants and products at equilibrium.

Magnitude of *K*

The larger the value of the numerator (concentration of the products), the larger the *K* value. The larger the value of the denominator (concentration of the reactants), the smaller the *K* value.

- If K is greater than 10^4, which is a large number, there is significantly more product(s) than reactant(s) at equilibrium.
- If K is less than 10^{-4}, which is a small number, there is significantly more reactant(s) than product(s) at equilibrium.
- If K is anywhere between these values, there is a significant amount of both reactants and products in the system at equilibrium.

See how to calculate K in Worked examples 6.5A and 6.5B.

6.5A Worked example
Find me in your obook pro

6.5A Worked example
Video demonstration

6.5B WORKED EXAMPLE

CALCULATING *K* FOR A REACTION FROM CONCENTRATIONS

Calculate the equilibrium constant, K, for the following equilibrium equation where the concentration of all reactants and products is 2 M:

$$2NO(g) + 2H_2(g) \rightleftharpoons N_2(g) + 2H_2O(g)$$

Solution

Think	Do
Step 1: Write an expression for K. Place the concentrations of the products in the numerator and the concentrations of the reactants in the denominator. Include the powers of the coefficients next to each concentration.	$K = \frac{[N_2]^1[H_2O]^2}{[NO]^2[H_2]^2}$
Step 2: Determine the units of K. If the coefficient is 2, include two Ms. If there is no coefficient, include one M.	$K = \frac{M \times (M \times M)}{(M \times M) \times (M \times M)}$ $= \frac{1}{M}$ $= M^{-1}$
Step 3: Substitute the values into the expression to calculate K. Remember to include the units.	$K = \frac{2 \times 2^2}{2^2 \times 2^2}$ $= \frac{2 \times 4}{4 \times 4}$ $= \frac{8}{16}$ $= 0.5\,M^{-1}$

Study tip

When you are asked to write an equilibrium expression for a reaction, it is not enough to show the expression alone (e.g. $\frac{[N_2][H_2O]^2}{[NO]^2[H_2]^2}$). You must include the equals sign with either K or the value of the constant (e.g. $K = \frac{[N_2][H_2O]^2}{[NO]^2[H_2]^2}$ or $\frac{[N_2][H_2O]^2}{[NO]^2[H_2]^2} = 0.5\,M^{-1}$).

Effects of changes on the *K* value

The only change that affects the K value of a system is temperature. This is because temperature affects the rates of the forward and reverse reactions. The rates of opposing reactions are rarely the same because of the difference in activation energy, so K will be different.

Additionally, when an equilibrium system partially opposes a change in temperature, it is opposing a change in energy and not concentration. The concentration fraction (amount of product and reactant) changes and, therefore, the K value also changes. Have a look at how to predict the effect of changing temperature on K in Worked example 6.5C. The K value can also change if the reaction itself is reversed or the coefficients change:

- Reversing a chemical equation will result in an inverse of the K value (e.g., if $K = 4$ and the reaction is reversed, $K = \frac{1}{4}$).
- Doubling the coefficients will square the K value (e.g., if $K = 4$ and the coefficients are doubled, $K = 4^2 = 16$).
- Halving the coefficients will result in the square root of the K value (e.g., if $K = 4$ and the coefficients are halved, $K = \sqrt{4} = 2$).

6.5C Worked example
Find me in your obook pro

6.5C Worked example
Video demonstration

Study tip

For endothermic reactions (positive ΔH), increasing temperature will increase K. For exothermic reactions (negative ΔH), increasing temperature will decrease K.

6.5 CHALLENGE

Calculating *K* for an oxidation reaction

The oxidation reaction of sulfur dioxide to form sulfur trioxide is shown below.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

4.0 mol of SO_2 and 2.8 mol of O_2 are reacted in a 6.0 L vessel and allowed to reach equilibrium at SLC. At this point, the percentage yield is 25%. Calculate the equilibrium constant for the reaction.

Using RICE tables to calculate concentrations

For systems where concentrations at equilibrium must be calculated using stoichiometry, a RICE table can be used.

R: ratio of coefficients

I: initial concentration (in M)

C: change that occurs (in M)

E: equilibrium concentration

Worked example 6.5D shows you how a RICE table is used to calculate concentrations of the different species in the reaction.

6.5D Worked example
Find me in your obook pro

6.5D Worked example
Video demonstration

6.5 CHECK YOUR LEARNING

Describe and explain

1 Describe the effect on the K value when the temperature of an exothermic reaction increases.

2 Describe the effect on the K value when the temperature of an endothermic reaction increases.

Apply, analyse and compare

3 Write an equilibrium expression for the following reactions.

  a $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

  b $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

  c $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

  d $CH_4(g) + H_2O(l) \rightleftharpoons CO(g) + 3H_2(g)$

4 State the units for the equilibrium reactions in Question 3.

5 Determine whether the K value will increase or decrease if the temperature was raised in the following equilibrium reaction systems:

  a $CO_2(g) + 2H_2O \rightleftharpoons CH_4(g) + 2O_2(g)$ $\Delta H = +802\ \text{kJ mol}^{-1}$

  b $N_2(g) + 3H_2 \rightleftharpoons 2NH_3(g)$ $\Delta H = -92\ \text{kJ mol}^{-1}$

  c $2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$ $\Delta H = -202\ \text{kJ mol}^{-1}$

6 N_2 and H_2 are pumped into a 4.0 L vessel to form NH_3. Initially, 1.0 mol of each is pumped into the container. At equilibrium, the concentration of N_2 is 0.20 M. Calculate the equilibrium constant (K).

7 Gaseous propane (C_3H_8) is combusted in a sealed 20 L container. If 2 M of propane and 13 M of oxygen are put into the container, and 150 mol of water is found at equilibrium, calculate the equilibrium constant (K).

Design and discuss

8 Discuss why the units of equilibrium constants are not always the same.

9 Discuss what an equilibrium constant represents. Your response must include an explanation of what the size of an equilibrium constant means in terms of the position of an equilibrium reaction.

6.6 The reaction quotient (*Q*)

KEY IDEAS

In this topic, you will learn that:

- the reaction quotient, *Q*, indicates whether a reaction is favouring reactants, products or an equal amount of both when it is not at equilibrium.

reaction quotient (*Q*) the ratio of reactants to products in a reaction when it is not at equilibrium

Systems that are not at equilibrium can still use the equilibrium law, but it instead calculates the **reaction quotient**, represented by *Q*. It is calculated in the same way, using the same equation as *K*:

$$Q = \frac{[C]^y[D]^z}{[A]^w[B]^x}$$

where the concentrations of species A, B, C and D are not necessarily at equilibrium.

Q provides an understanding of the amount of reactants and products in a reaction system *before* it reaches equilibrium. With this, an experimenter can identify the direction in which the reaction must proceed to reach equilibrium.

- If *Q* is greater than *K*, there are more products than reactants compared to when the system is at equilibrium. There will be a net reverse reaction to form more reactants and reach equilibrium.
- If *Q* is smaller than *K*, there are more reactants than products compared to when the system is at equilibrium. There will be a net forward reaction to form more products and reach equilibrium.

6.6 SKILL DRILL

Designing an experiment to optimise the formation of products

Key science skill: Plan and conduct investigations

A pharmaceutical scientist is tasked with the job of optimising the production of a medicine. They find that the reaction to form the final product is an equilibrium reaction.

FIGURE 1 Pharmaceutical research

Practise your skills

1 Describe an experiment the pharmaceutical scientist could conduct to determine the factors that increase the yield of the reaction. In your answer, indicate how they could check whether the reaction has reached equilibrium.

2 Discuss the significance of the finding that the reaction is an equilibrium reaction. How does this help the optimisation of processes to make the medicine?

Need help planning experiments? See Topic 1.4 (page 14).

6.6 WORKED EXAMPLE

IDENTIFYING THE DIRECTION IN WHICH A REACTION MUST PROCEED TO REACH EQUILIBRIUM

In a gaseous system, 36 mol of H_2, 16 mol of N_2 and 28 mol of NH_3 are initially pumped into a 4 L container. At equilibrium, the concentration of hydrogen is measured to be 3 M.

a Calculate the reaction quotient (Q) for the reaction when the chemicals are initially added to the vessel.

b Calculate the equilibrium constant (K).

c Compare the Q and K values and determine in which direction the reaction has moved to reach equilibrium.

Solution

<table>
<tr><th>Think</th><th>Do</th></tr>
<tr><td>Step 1: Write a balanced chemical equation for the reaction.</td><td>$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$</td></tr>
<tr><td>Step 2: Convert all amounts, in mol, to concentration using $c = \frac{n}{V}$. Remember that there are no products in the container initially.</td><td><table><tr><td>$[N_2] = \frac{n}{V}$
$= \frac{16}{4}$
$= 4\,M$</td><td>$[H_2] = \frac{n}{V}$
$= \frac{36}{4}$
$= 9\,M$</td><td>$[NH_3] = \frac{n}{V}$
$= \frac{28}{4}$
$= 7\,M$</td></tr></table></td></tr>
<tr><td>Step 3: Substitute the concentration values into the equation to calculate Q.</td><td>a $Q = \frac{[NH_3]^2}{[H_2]^3[N_2]}$
$= \frac{7^2}{9^3 \times 4}$
$= \frac{49}{2916}$
$= 0.0168\,M^{-2}$
The value is small, so there are more reactants present in the system than products.</td></tr>
<tr><td>Step 4: Fill in the RICE table to calculate the concentrations at equilibrium. See Worked example 6.5D if you need a refresher!</td><td><table><tr><th></th><th>N_2</th><th>H_2</th><th>NH_3</th></tr><tr><td>R</td><td>1</td><td>3</td><td>2</td></tr><tr><td>I</td><td>4</td><td>9</td><td>7</td></tr><tr><td>C</td><td>$N_2 : NH_3$ is
1 : 2
$\frac{1}{2} \times 4 = 2$</td><td>$H_2 : NH_3$ is
3 : 2
$\frac{3}{2} \times 4 = 6$</td><td>+4</td></tr><tr><td>E</td><td>4 – 2 = 2</td><td>9 – 6 = 3</td><td>7 + 4 = 11</td></tr></table></td></tr>
<tr><td>Step 5: Substitute the concentrations into the equilibrium expression to calculate K.</td><td>b $K = \frac{[NH_3]^2}{[H_2]^3[N_2]}$
$= \frac{11^2}{3^3 \times 2}$
$= \frac{121}{54}$
$= 2.24\,M^{-2}$
At equilibrium, the system has similar amounts of reactants and products.</td></tr>
<tr><td>Step 6: Compare the Q and K values to determine the direction that the reaction has moved in.</td><td>c $Q < K$, so after addition of the chemicals to the container, there was a net forward reaction to reach equilibrium.</td></tr>
</table>

Study tip

Assumptions are made in equilibrium. Be careful if a question says that the container was initially empty or vacated. This means that if only reactants are placed into the vessel at time 0, there are no products, so their initial concentration is 0.

6.6 CHECK YOUR LEARNING

Describe and explain

1 Explain the difference between the equilibrium constant and the reaction quotient.

2 Explain when the reaction quotient might be calculated instead of the equilibrium constant.

3 Explain what you can deduce about a reaction if its equilibrium constant and reaction quotient are the same.

Apply, analyse and compare

4 Explain what direction is predominant in a reaction if its Q value is larger than its K value.

5 Explain what direction is predominant in a reaction if its Q value is smaller than its K value.

6 Explain the assumptions that can be made about the concentration of reactants and products in a reaction if their Q values are:

- **a** 9.2×10^{-16} M
- **b** 0.12 M
- **c** 3.5×10^{9} M.

7 10.0 mol of SO_2 and 6.0 mol of O_2 are added to a 2.0 L vessel. Thirty seconds later the concentration of SO_3 is measured at 4 M.

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $K = 2.0 \times 10^{-2}\,M^{-1}$

- **a** Calculate the reaction quotient, Q.
- **b** Communicate in which direction the reaction must proceed to reach equilibrium and justify your answer using the K and Q values.

8 Initially, 6.0 M NO and 8.0 M O_2 are placed in a 1.5 L vessel. 10 seconds into the reaction, the concentration of NO was found to have decreased by 6 mol.

$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$ $K = 2.00\,M^{-1}$

Evaluate the above information and answer the following questions.

- **a** Calculate the reaction quotient, Q.
- **b** Communicate in which direction the reaction must proceed to reach equilibrium and justify your answer using the K and Q values.

6.7

Green chemistry and equilibrium

KEY IDEAS

In this topic, you will learn that:

- there is a trade-off between the optimal rate and temperature
- green chemistry principles can be applied by using catalysts and designing reactions to be energy efficient.

In this chapter, you have learned that we can make many different changes to an equilibrium system to increase the rate of a reaction or yield of the desired product. In some cases, there is conflict between these two outcomes. For example, changing the temperature to increase the rate of a reaction might actually decrease the quantity of products formed.

In addition, optimisation of processes may generate waste products and use up a lot of energy, which contradicts the green chemistry principles. Since 1998, when the framework was first developed for making a greener chemical process or product, industry has been improving processes to decrease the amount of energy used and the amount of waste produced.

In this topic, you will explore how the green chemistry principles play a role in the optimisation of equilibrium reactions. You will then see how these principles are applied to address the conflict between optimal rate and temperature for the Haber process, where nitrogen and hydrogen gases react in the presence of a catalyst to form ammonia.

Green chemistry principles for equilibrium reactions

In Chapter 5, you learnt that increasing temperature can increase the rate of a reaction. But you might also remember, from this chapter, that increasing temperature may not favour the formation of product.

Recall that when you add energy as a product in an exothermic reaction, the position of equilibrium will shift to partially oppose this change by causing a net reverse reaction. This means that the yield of desired product decreases (Figure 1).

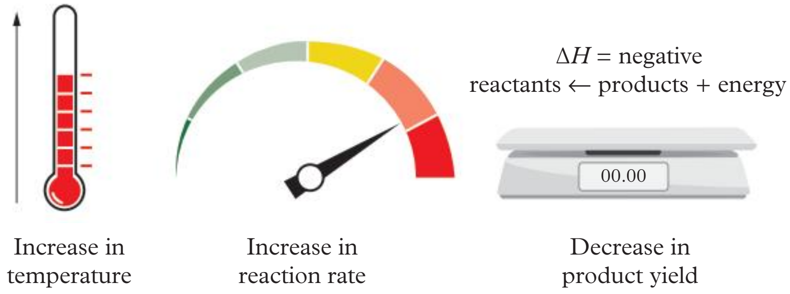


FIGURE 1 Increasing temperature can decrease the yield of products.

Luckily, we can do more than just increase the temperature to increase the reaction rate. We can optimise equilibrium reactions to address this conflict, and also do this in a way that is 'green': by making sure to be efficient with energy usage and by using catalysts where possible.

6.7A
PRACTICAL: SIMULATION

What factors affect the Haber process? Go to your **Student obook pro**.

6.7B
PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT

How can the formation of sulfuric acid be optimised? Go to your **Student obook pro**.

Design for energy efficiency

Increasing energy efficiency involves running chemical reactions at room temperature and pressure whenever possible. Reactions that require extreme temperatures (higher or lower than room temperature) to increase the rate or yield of the reaction must obtain this heat energy from electricity. Changing the pressure of a reaction system also requires electricity.

In Victoria, much of the electricity obtained from the grid comes from the state's brown coal deposits. Here, the coal is combusted in thermal power stations to generate electricity, which is then used to heat or cool reactions. This process wastes a significant amount of energy as heat, which makes it inefficient. More heat is then lost when we use electricity to heat the chemical reaction.

The additional energy requirement of these reactions makes them less efficient. The focus of this green chemistry principle is to design processes that only require room temperature and atmospheric pressure.

FIGURE 2 Much of the electricity in Victoria is sourced from coal.

Catalysis

The principle of catalysis states that 'catalysts should be selected to generate the same desired product(s) with less waste and using less energy and reagents in reaction processes/pathways'. Many of these catalysts help to accelerate reactions at room temperature and atmospheric pressure, shifting us away from relying on temperature and pressure changes.

Waste is also minimised by using catalysts in reactions because catalysts are effective in small amounts, can be reused and can perform the same catalytic role in the same reaction many times over. As they do not participate in and are therefore not consumed by the reaction, catalysts are preferred over adding reactants (or reagents) to the system.

The importance of catalysts is evident in the chemical industry: approximately 90% of industrial chemical processes use some form of catalytic reactions. There have also been at least 15 Nobel prizes awarded for research into catalytic processes. The advantages and disadvantages of using catalysts are summarised in Table 1.

Study tip

Remember that catalysts lower the activation energy of a reaction by providing an alternative reaction pathway for both forward and reverse reactions.

TABLE 1 Advantages and disadvantages of using catalysts

Advantages	Disadvantages
• Increase the rate of a chemical reaction • Reduce the need for excess temperatures, which have high energy costs • Can be reused • Decrease the need for excess reactants, also potentially reducing the generation of waste (e.g. excess reactants, undesired products)	• Do not increase the yield of a reaction • May still require changes based on Le Chatelier's principle to increase yield • May themselves be harmful or toxic to the environment • Can lose their activity after a certain period of use

FIGURE 3 Iron oxide is used as a catalyst.

Optimising the Haber process

The Haber process, designed for chemical warfare in the early 1900s, has since been useful to generate ammonia to fertilise crops for the mass production of food. The process involves first extracting nitrogen gas from the atmosphere and generating hydrogen gas from a separate reaction.

The formation of hydrogen

To generate the hydrogen required in the reaction, methane is reacted with water vapour according to the chemical equation:

$$CH_4(g) + H_2O(l) \rightleftharpoons CO(g) + 3H_2(g) \qquad \Delta H = +214.68\,kJ\,mol^{-1} \qquad K = 1.24 \times 10^{-25}\,M^2$$

Table 2 summarises the ways in which reaction conditions can be changed, and their effects on reaction rate and yield. You can see that there are a number of compromises when we add chemical species and change pressure or volume. While increasing temperature seems to improve reaction rate and yield, this comes with a high energy and environmental cost. These can be overcome by adding a catalyst (usually nickel).

TABLE 2 Conditions for increasing the rate of reaction and yield of the hydrogen

Change	Increases rate	Increases yield	Things to consider
Adding reactants	✓	✓	Will increase the amount of waste and pressure in a fixed-volume system
Increasing temperature	✓	✓	Causes a net forward reaction and increases yield
Increasing pressure/ decreasing volume	✓		Safety issues associated with high-pressure systems – excess pressure can cause weak reaction vessels to break open and release harmful chemicals
Decreasing pressure/ increasing volume		✓	Reduces the concentration of reactants and frequency of collisions
Adding a catalyst	✓		Waste management of Ni(s) catalysts and environmental toxicity

The formation of ammonia

Nitrogen gas and hydrogen gas are then reacted to form ammonia, according to the following chemical equation:

$$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$$
$$\Delta H = -92\,kJ\,mol^{-1} \quad K = 5.95 \times 10^5\,M^{-2}$$

Table 3 summarises the ways in which reaction conditions can be changed, and their effects on reaction rate and yield. You can see that there are a number of compromises when we add chemical species, change temperature, and change pressure or volume. These can be overcome by adding a catalyst (usually iron or iron oxide).

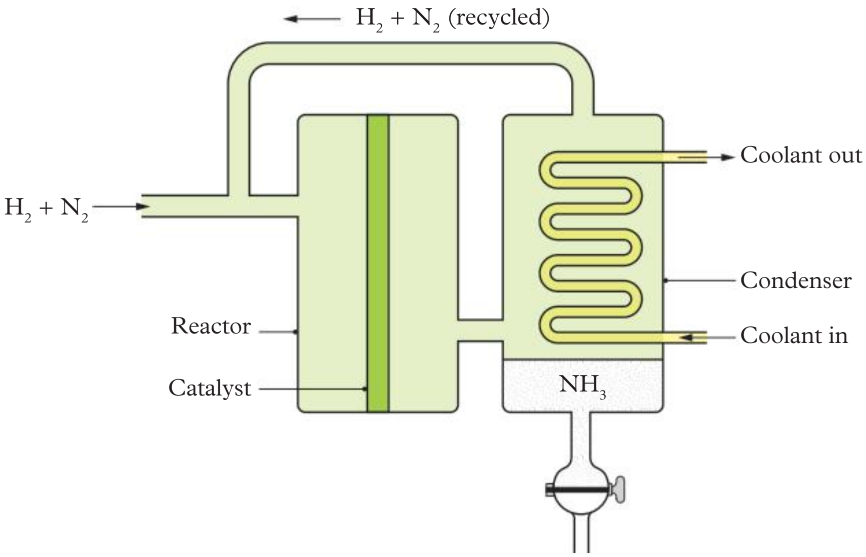


FIGURE 4 The Haber process for producing ammonia

TABLE 3 Conditions for increasing the rate of reaction and yield of ammonia

Change	Increases rate	Increases yield	Things to consider
Adding reactants	✓	✓	Will increase the amount of waste
Increasing temperature	✓		Causes a net reverse reaction and decreases yield
Decreasing temperature		✓	Reduces the rate of reaction
Increasing pressure/ decreasing volume	✓	✓	Safety issues associated with high-pressure systems – excess pressure can cause weak reaction vessels to break open and release harmful chemicals
Adding a catalyst	✓		

While the use of catalysts has its own disadvantages, they seem to be outweighed by the benefits.

6.7 CHECK YOUR LEARNING

Describe and explain

1 Explain the issue with the impact of increasing temperature of the rate and yield of exothermic reactions.

2 Explain the issue with the impact of increasing pressure on the rate and yield of some reactions.

3 Explain the effects of changing reaction stoichiometry and using catalysts on the reaction yield.

Apply, analyse and compare

4 Explain how the rate and yield of the following exothermic chemical reactions can be optimised.

a $2C_4H_{10}(g) + 18O_2(g) \rightleftharpoons 8CO_2(g) + 10H_2O(l)$ $\Delta H = -2880\,kJ\,mol^{-1}$

b $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$ $\Delta H = -184.60\,kJ\,mol^{-1}$

c $2H_2S(g) + 3O_2(g) \rightleftharpoons 2H_2O(l) + 2SO_2(g)$ $\Delta H = -1036.96\,kJ\,mol^{-1}$

5 Explain how the rate and yield of the following chemical reactions can be optimised.

a $2CH_4(g) + S_8(g) \rightleftharpoons 2CS_2(g) + 4H_2S(g)$ $\Delta H = +201.91\,kJ\,mol^{-1}$

b $Sb_2S_3(s) + Fe(s) \rightleftharpoons Sb(s) + FeS(s)$ $\Delta H = +74.89\,kJ\,mol^{-1}$

c $N_2(g) + 2O_2(g) \rightleftharpoons 2NO_2(g)$ $\Delta H = +66.36\,kJ\,mol^{-1}$

Design and discuss

6 Discuss the green chemistry principle of catalysis and explain the rate and yield advantages of using catalysts in equilibrium reactions.

7 Discuss the green chemistry principle of energy efficiency and explain how equilibrium reactions can be more energy efficient while maximising rate and yield.

8 Explain how the rate and equilibrium trade-off applies to the production of ammonia:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92\,kJ\,mol^{-1}$

Investigate the Haber process and communicate how the trade-off is overcome, including the temperature and pressure conditions and how these maximise the yield of ammonia.

CHAPTER 6 Review

Chapter summary

6.1

- Reversible reactions can proceed in forward and reverse directions. Irreversible reactions can only proceed in the forward direction.
- The rate of a reaction is the measure of how quickly the forward reaction occurs. The extent of a reaction is an indication of its magnitude as it moves in a forward direction.

6.2

- Equilibrium reactions are reversible. Dynamic equilibrium is reached when the rate of the forward and reverse reaction is the same.
- Homogeneous equilibrium applies to systems in which all reactants and products are in the same state.

6.3

- Le Chatelier's principle states that equilibrium systems will partially oppose any change made to them. It explains the effect that changes in addition or removal of species, volume, pressure and temperature have on the position of a reaction and can be summarised as having resulted in a net forward reaction, net reverse reaction, or as having no effect on the reaction system.

6.4

- The yield of a reaction is a measure of the amount of desired product. Le Chatelier's principle can be applied to maximise yield, by making changes to cause a net forward reaction.

6.5

- The equilibrium law outlines a mathematical relationship between the concentration of reactants and products in an equilibrium reaction. The equilibrium constant (K) represents this ratio at equilibrium.
- K can only be changed by an increase or decrease in temperature.
- Concentrations at equilibrium can be calculated using a RICE table.

6.6

- For reactions not at equilibrium, the equilibrium law calculates a reaction quotient (Q).
- A RICE table can be used to calculate concentrations of reactants and products in a reaction system at any time. Q values can be calculated from these concentrations and compared to K values to determine the direction in which a reaction must proceed to reach equilibrium.

6.7

- Increasing the temperature increases the reaction rate, but can decrease the yield of products, especially for exothermic reactions.
- Catalysts increase the rate of a reaction but not the yield. Therefore, they have no effect on the equilibrium, but will allow a reaction to reach equilibrium at a faster rate.
- The green chemistry principles of catalysis and the design for energy efficiency are based on the addition of a catalyst to a chemical reaction to lower the activation energy of the reaction and ensure that excess energy is not wasted to increase the rate of a chemical reaction. They also ensure that minimal amounts of chemicals are wasted.

Key formulas

Concentration	$c = \frac{n}{V}$
Equilibrium expression	$K = \frac{[C]^y[D]^z}{[A]^w[B]^x}$
Reaction quotient	$Q = \frac{[C]^y[D]^z}{[A]^w[B]^x}$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Mostly	Not really	Revision link
• distinguish between reversible and irreversible reactions, and between the rate and extent of a reaction	☐	☐	☐	Go back to Topic 6.1
• describe the dynamic nature of homogeneous equilibria involving aqueous solutions or gases	☐	☐	☐	Go back to Topic 6.2
• represent equilibria by balanced chemical or thermochemical equations (including states) and by concentration–time graphs	☐	☐	☐	Go back to Topic 6.2
• explain how the position of equilibrium changes when changes in temperature or species or volume (concentration or pressure) are applied to a system at equilibrium	☐	☐	☐	Go back to Topic 6.3
• represent changes to the position of equilibrium using concentration–time graphs	☐	☐	☐	Go back to Topic 6.3
• apply Le Chatelier's principle to identify factors that favour the yield of a chemical reaction	☐	☐	☐	Go back to Topic 6.4
• use equilibrium expressions to calculate the equilibrium constant (*K*, including units) for a closed homogeneous equilibrium system	☐	☐	☐	Go back to Topic 6.5
• explain how system temperature affects the equilibrium constant (*K*)	☐	☐	☐	Go back to Topic 6.5
• use the reaction quotient (*Q*) as a quantitative measure of the extent of a chemical reaction; that is, the relative amounts of products and reactants present during a reaction at a given point in time	☐	☐	☐	Go back to Topic 6.6
• describe responses to the conflict between optimal rate and temperature considerations in producing equilibrium reaction products, with reference to the green chemistry principles of catalysis and designing for energy efficiency	☐	☐	☐	Go back to Topic 6.7

Revision questions

Multiple choice

1 Consider the following reaction.

$$CH_4(g) + H_2O(l) \rightleftharpoons CO(g) + 3H_2(g) \quad \Delta H = +206\,\text{kJ mol}^{-1}$$

Which changes would maximise the yield of hydrogen?

A Decreasing volume and decreasing temperature

B Decreasing volume and increasing temperature

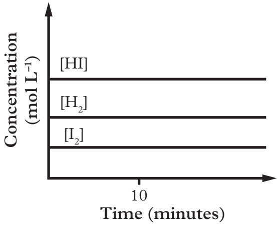
C Increasing volume and decreasing temperature

D Increasing volume and increasing temperature

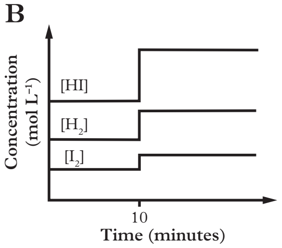
2 Which graph represents the effect of a decrease in volume on the concentration of reactants and products in the following reaction?

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g)$$

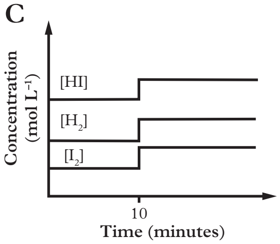
A



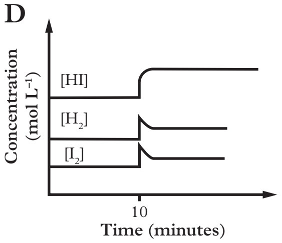
B



C



D



3 Gaseous NOCl decomposes to form the gases NO and Cl_2 according to the following equation.

$$2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g) \quad K = 1.6 \times 10^{-5}\,\text{M}$$

What is the numerical value of the equilibrium constant for the following reaction?

$$NO(g) + \tfrac{1}{2}Cl_2(g) \rightleftharpoons NOCl(g)$$

A -1.6×10^{-5}

B 1.6×10^{-5}

C 2.5×10^{2}

D 6.3×10^{4}

4 Which of the following statements regarding dynamic equilibrium is false?

A Reactants and products are breaking down and re-forming in the molar ratio.

B Both reactants and products are present in the reaction.

C Molar amounts of both reactants and products are equal.

D The rate of the forward and reverse reactions is the same.

Use the following equation to answer Questions 5–8.

Butane is allowed to combust in oxygen in a sealed container according to the following equation:

$$2C_4H_{10}(g) + 13O_2(g) \rightleftharpoons 8CO_2(g) + 10H_2O(g) \quad \Delta H = -\text{ve}$$

5 The equilibrium expression used to represent the equation is:

A $K = \dfrac{[C_4H_{10}]^2 \times [O_2]^{13}}{[CO_2]^8 \times [H_2O]^{10}}$

B $K = \dfrac{8[CO_2] \times 10[H_2O]}{2[C_4H_{10}] \times 13[O_2]}$

C $K = \dfrac{2[C_4H_{10}] \times 13[O_2]}{8[CO_2] \times 10[H_2O]}$

D $K = \dfrac{[CO_2]^8 \times [H_2O]^{10}}{[C_4H_{10}]^2 \times [O_2]^{13}}$

6 To maximise the yield, a chemist should:

A increase the concentration of products.

B add an inert gas.

C double the volume.

D increase the pressure.

7 To maximise the concentration of butane:

A increase the temperature and increase the volume.

B increase the temperature and decrease the volume.

C decrease the temperature and decrease the volume.

D decrease the temperature and increase the volume.

8 The effect on the K value of the equation when the reaction is reversed and the coefficients are doubled is:

A $\frac{1}{K^2}$

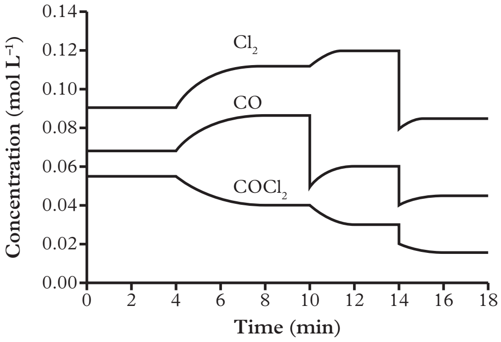
B $\frac{1}{2 \times K}$

C K^2

D $\sqrt{K^2}$

Use the following information to answer Questions 9 and 10.

$COCl_2(g) \rightleftharpoons Cl_2(g) + CO(g) \quad \Delta H = +108\,kJ$



9 What change has been made to the system at the 4-minute mark?

A There is an increase in volume.

B There is an increase in temperature.

C There is a decrease in temperature.

D There is a decrease in volume.

10 What change has been made to the system at the 14-minute mark?

A There is an increase in volume.

B An inert gas has been added.

C A catalyst has been added.

D There is a decrease in volume.

Short answer

Describe and explain

11 Write the equilibrium expression for the following reactions:

a $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

b $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$

12 Explain the impact that the addition of a catalyst has on an equilibrium system.

13 Explain the effect of adding an inert gas into an equilibrium system.

14 An equilibrium system contains a mixture of colourless N_2O_4 and dark brown NO_2. At 25°C, the equilibrium constant for the reaction is 4.65×10^{-3}.

Initially the equilibrium system is a light brown colour.

a Use an equilibrium expression to explain the colour of the system at 25°C.

b Explain the colour changes observed and their intensities as the volume of the system doubles.

c Explain the colour changes observed and their intensities as the temperature of the system increases (ΔH = +ve).

d Some colourless inert gas is injected into the system. Explain the colour changes observed and their intensities as the gas is added.

15 Consider the following equilibrium system:

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons FeSCN^{2+}(aq)$$

Discuss the effect of the following changes and use Le Chatelier's principle to justify the effect on the position of equilibrium.

a $AgNO_3(aq)$ is added to the reaction.

b NaF(aq) is added to the reaction.

c The reaction is diluted.

d NaOH is added to the reaction.

e $Pb(NO_3)_2$ is added to the reaction.

16 Consider the following equation:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

At 20°C, the equilibrium constant is 1.3×10^{-2} and at 400°C it is 25.6. Explain whether the reaction is exothermic or endothermic and justify your answer with the data given.

17 Methanol can be formed using the following reaction:

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

$$\Delta H = +128\,kJ\,mol^{-1}$$

At 25°C the equilibrium constant for the reaction is measured as $3.9 \times 10^{-2}\,M^{-2}$.

a Explain what can be concluded about the concentration of reactants and products.

b Use Le Chatelier's principle to explain the effect of:

  - **i** adding methanol to the reaction vessel at a constant temperature
  - **ii** decreasing the temperature of the system
  - **iii** doubling the volume of the system.

18 An equilibrium system has a larger *Q* value than *K* value. Explain what must happen for the reaction to reach equilibrium.

19 An equilibrium system has a smaller *Q* value than *K* value. Explain what must happen for the reaction to reach equilibrium.

20 When a change is made to an equilibrium system, explain why it is essential to use the terminology *partially oppose*, rather than just saying that an equilibrium system opposes a change.

Apply, analyse and compare

21 Nitrosyl chloride (NOCl) decomposes according to the equation:

$$2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$$

Initially, 3.6 mol of NOCl(g) is placed in an empty 2.5 L flask and allowed to reach equilibrium. If the concentration of Cl_2 is 0.50 M at equilibrium, calculate the equilibrium constant for the reaction.

22 Consider the following equilibrium equations:

$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) \qquad K = 4.3 \times 10^{-25}$

$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g) \qquad K = 6.4 \times 10^{9}$

- **a** Identify the units for the equilibrium constants in the two reactions.
- **b** Calculate the equilibrium constant for the following reactions.
  - **i** $2NO(g) \rightleftharpoons N_2(g) + O_2(g)$
  - **ii** $4NO(g) + 2O_2(g) \rightleftharpoons 4NO_2(g)$
  - **iii** $\frac{1}{3}N_2(g) + \frac{1}{3}O_2(g) \rightleftharpoons \frac{2}{3}NO(g)$
  - **iv** $NO(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO_2(g)$
  - **v** $\frac{1}{2}NO_2(g) \rightleftharpoons \frac{1}{2}NO(g) + \frac{1}{4}O_2(g)$

23 Hydrogen iodide is formed according to the following equilibrium reaction.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

Initially, 0.90 mol of H_2 and 0.50 mol of I_2 are added to a 2.00 L vessel. Two minutes into the reaction, the concentration of HI is measured at 0.40 M.

- **a** Calculate *Q*.
- **b** The reaction has a *K* value of 0.500 at equilibrium. Using the *Q* and *K* values, justify the direction in which the reaction must move to reach equilibrium.
- **c** Identify the units of the equilibrium constant. Show your working.
- **d** The volume of the system at equilibrium is halved. Explain the effect of this change on the equilibrium system.
- **e** Create a concentration–time graph to demonstrate the effect of halving the volume of the system from part **d**.

24 Consider the equilibrium mixture between complex copper ions below:

$$\left[Cu(H_2O)_4\right]^{2+}(aq) + 4Cl^-(aq) \rightleftharpoons \left[CuCl_4\right]^{2-}(aq) + 4H_2O(l)$$

In the mixture, $Cu(H_2O)_4^{2+}$ is a dark blue colour and $CuCl_4^{2-}$ is yellow. When cooled in an ice bath, the mixture becomes darker blue, but in a hot water bath, it turns green.

- **a** Identify whether the reaction is exothermic or endothermic. Use the colour changes in the water baths to justify your answer.
- **b** Draw a reaction profile for the reverse reaction.
- **c** A solution of $Pb(NO_3)_2(aq)$ is added to the reaction mixture. Create a concentration–time graph representing the change made and the system after the change.

25 Initially, 1.50 mol of $NO_2(g)$ is pumped into an empty 4.00 L container and allowed to reach equilibrium. The NO_2 will decompose into NO and O_2 gases according to the following equation.

$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g) \qquad K = 6.4 \times 10^{9}$

- **a** Explain why it can be assumed that the amount in mol of $NO_2(g)$, at equilibrium, is still 1.50 mol.
- **b** Calculate the concentration of NO(g) at equilibrium.
- **c** Calculate the concentration of $O_2(g)$ at equilibrium.

26 Initially, 2.0 mol of HI(g) is pumped into an empty 2.00 L container and allowed to reach equilibrium. The HI will decompose into H_2 and I_2 gases according to the following equation:

$2HI(g) \rightleftharpoons H_2(g) + I_2(g) \qquad K = 5.6 \times 10^{-5}$

a Explain why it can be assumed that the amount of HI at equilibrium is still 2.0 mol.

b Calculate the concentration of $H_2(g)$ at equilibrium.

c Calculate the concentration of $I_2(g)$ at equilibrium.

Design and discuss

27 Investigate the concept of a buffer and communicate how it is able to maintain the pH of equilibrium systems. Use human blood and its ability to regulate CO_2 as an example to justify your answer.

28 In a reaction between hydrogen and nitrogen gases, ammonia is formed according to the following equation:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

$\Delta H = -92.4\ kJ\ mol^{-1}$

$K = 5.95 \times 10^5\ M^{-2}$ at 298 K

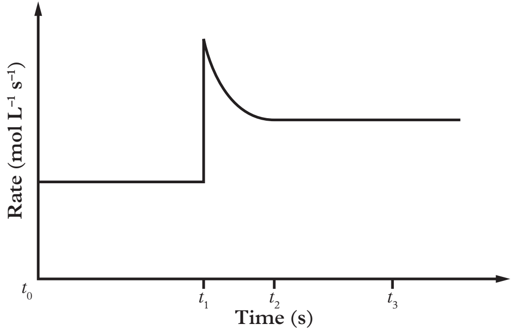
a Write an equilibrium expression for the reaction.

Initially, 3.0 mol each of hydrogen and nitrogen gases are pumped into an empty 2.0 L container.

b Calculate the reaction quotient if the concentration of ammonia is 0.50 M.

c Explain the direction that the reaction must favour to reach equilibrium.

29 The graph below demonstrates the rate of formation of ammonia when a change is made at t_1.



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Use the rate–time graph above to explain what is occurring at the following time intervals:

a Between t_0 and t_1

b At t_1, when the volume of the container is decreased

c Between t_1 and t_2

d Between t_2 and t_3

30 Investigate the production of nitric acid.

a Write the three equilibrium reaction equations and state whether they are endothermic or exothermic.

b State the optimal conditions for the production of nitric acid for each equation (i.e. temperature, pressure, concentration, catalyst, etc.).

c Discuss where the trade-off exists in the production of nitric acid. Explain the concept of a trade-off and communicate how it is overcome in this process.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER

7

Production of chemicals using electrolysis

KEY KNOWLEDGE

- → the use and limitations of the electrochemical series to explain or predict the products of the electrolysis of particular chemicals, given their state (molten liquid or in aqueous solution) and the electrode materials used, including the writing of balanced equations (with states) for the reactions occurring at the anode and the cathode and the overall redox reaction for the cell
- → the common design features and general operating principles of commercial electrolytic cells (including, where practicable, the removal of products as they form), and the selection of suitable electrode materials, the electrolyte (including its state) and any chemical additives that result in a desired electrolysis product (no specific cell is required)
- → the common design features and general operating principles of rechargeable (secondary) cells with reference to discharging as a galvanic cell and recharging as an electrolytic cell, including the conditions required for the cell reactions to be reversed and the electrode polarities in each mode
- → the role of innovation in the designing cells to meet society's energy needs in terms of producing 'green' hydrogen (including equations in acidic conditions) using the following methods:
  - – polymer electrolyte membrane electrolysis powered by either photovoltaic (solar) or wind energy
  - – artificial photosynthesis using a water oxidation and proton reduction catalyst system
- → the application of Faraday's Laws and stoichiometry to determine the quantity of electrolytic reactant and product, and the current or time required to either use a particular quantity of reactant or produce a particular quantity of product.

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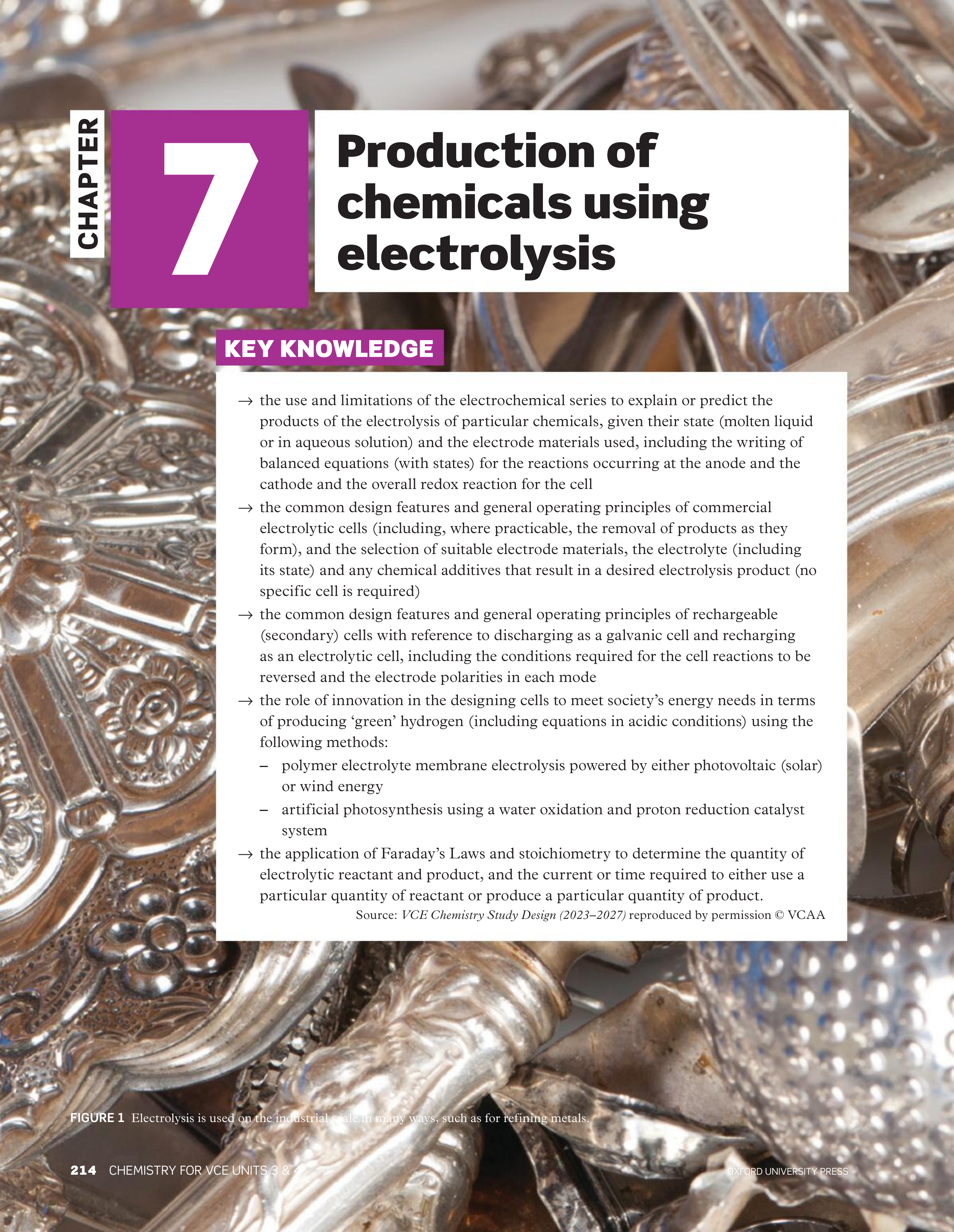
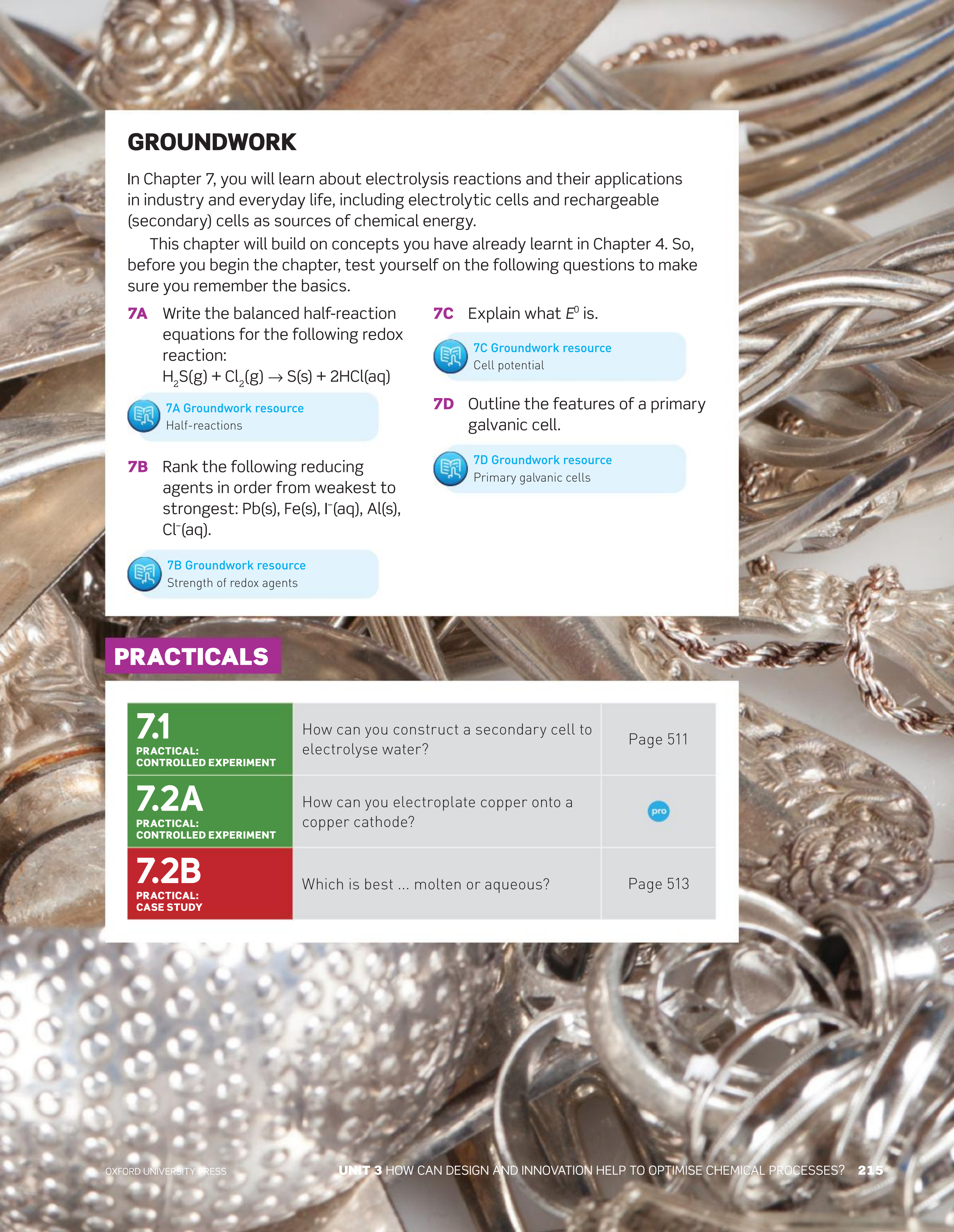


FIGURE 1 Electrolysis is used on the industrial scale in many ways, such as for refining metals.



GROUNDWORK

In Chapter 7, you will learn about electrolysis reactions and their applications in industry and everyday life, including electrolytic cells and rechargeable (secondary) cells as sources of chemical energy.

This chapter will build on concepts you have already learnt in Chapter 4. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

7A Write the balanced half-reaction equations for the following redox reaction:
$H_2S(g) + Cl_2(g) \rightarrow S(s) + 2HCl(aq)$

7A Groundwork resource
Half-reactions

7B Rank the following reducing agents in order from weakest to strongest: Pb(s), Fe(s), $I^-(aq)$, Al(s), $Cl^-(aq)$.

7B Groundwork resource
Strength of redox agents

7C Explain what E^0 is.

7C Groundwork resource
Cell potential

7D Outline the features of a primary galvanic cell.

7D Groundwork resource
Primary galvanic cells

PRACTICALS

7.1 **PRACTICAL: CONTROLLED EXPERIMENT**	How can you construct a secondary cell to electrolyse water?	Page 511
7.2A **PRACTICAL: CONTROLLED EXPERIMENT**	How can you electroplate copper onto a copper cathode?	pro
7.2B **PRACTICAL: CASE STUDY**	Which is best ... molten or aqueous?	Page 513

7.1 Electrolysis reactions

KEY IDEAS

In this topic, you will learn that:

- in electrolysis, electrical energy is used to reverse spontaneous redox reactions
- the products of electrolysis can be predicted using the electrochemical series if the molten, aqueous or electrode reactants are known.

In Chapter 4, you learnt that spontaneous redox reactions use chemical energy to produce electrical energy. Electrical energy can also be applied to a system to produce chemical energy. This reverse, non-spontaneous redox reaction is one of the leading chemical techniques for the separation and purification of compounds, especially metals.

In this topic, you will learn about the requirements for this non-spontaneous redox reaction and how to predict reaction products.

Electrolysis

electrolysis
the process by which electrical energy is passed into a cell, using a power source, resulting in the reversal of a spontaneous redox reaction

Electrolysis is a type of non-spontaneous redox reaction. An external electrical energy source is used to provide energy for the reaction to occur. Think back to Chapter 6: if a forward reaction (i.e. the spontaneous redox reaction) generates 5.0 V, then for the reverse reaction (i.e. the non-spontaneous redox reaction) to occur, more than 5.0 V needs to be supplied to the system.

Electron movement in electrolysis

Electrolysis reactions occur in an electrolytic cell, which has some similarities to a galvanic cell. In both, electrons are released from the anode and consumed at the cathode. The main difference between the two is that in electrolytic cells, voltage is applied to push the electrons towards the cathode.

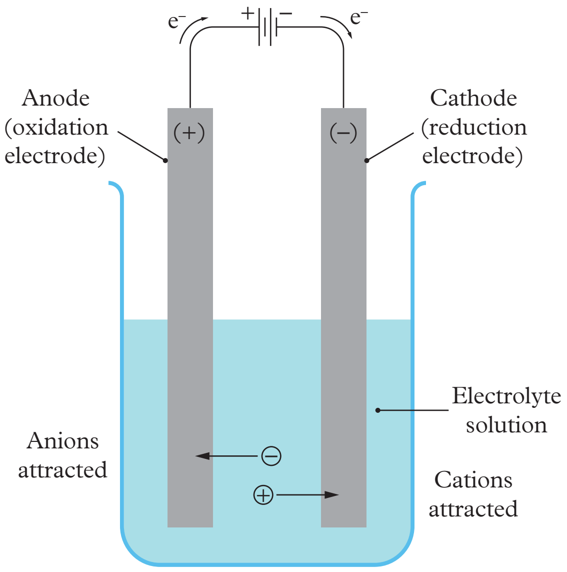


FIGURE 1 A basic diagram of an electrolytic cell

The anode is connected to the positive terminal of the power supply and becomes positive. Electrons are pulled from the electrolyte solution (or electrode itself) into and away from the anode. Negative ions in the electrolyte move towards the anode, where they lose electrons (oxidation):

$$X^- \rightarrow X + e^-$$

The cathode is connected to the negative terminal of the battery and becomes negative. Electrons are pushed onto the cathode, which can then supply electrons to the electrolyte solution. Positive ions move towards the negative cathode to gain electrons (reduction):

$$M^+ + e^- \rightarrow M$$

Study tip

In an electrolytic cell, oxidation occurs at the positive anode and reduction occurs at the negative cathode.

7.1
PRACTICAL: CONTROLLED EXPERIMENT

How can you construct a secondary cell to electrolyse water? Go to page **511**.

Predicting electrolysis reactions using the electrochemical series

Like with the spontaneous redox reactions in Chapter 4, the electrochemical series can be used to predict the products of non-spontaneous electrolysis reactions. This is possible as long as we know the state of the reactants: molten salts, aqueous electrolytes, or solid reactive electrodes.

We can use the **OX Z RED** method, which uses a diagram shaped as the letter 'Z' (Figure 2).

1 Find the appropriate half-equations on the electrochemical series.
2 Identify the chemical species that are reacting.
3 The strongest oxidising agent will react with the strongest reducing agent when sufficient electrical energy is applied. This is summarised in the 'Z' configuration of Figure 2.
4 **Ox**idation is on the top part of the **Z** and **red**uction is on the bottom, so the acronym is **OX Z RED**. Rewrite the half-equations with the reactants and products on the correct sides of the reaction arrow.

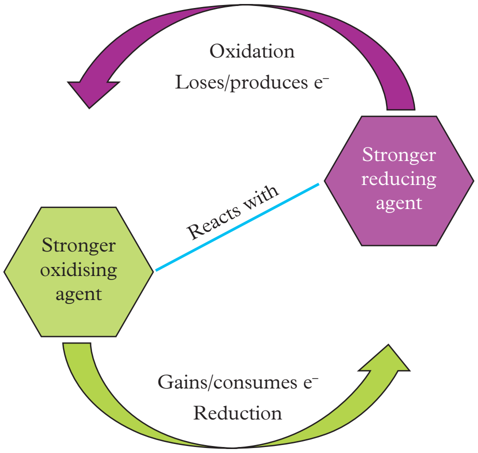


FIGURE 2 The 'Z' configuration is used to determine the non-spontaneous electrolysis reaction that can occur.

Molten electrolysis

When solid salts are melted at very high temperatures, they form **molten** salts. Their state becomes liquid (l). This is different from aqueous (aq) salts, which are dissolved in water. The electrolysis of molten salts must therefore occur at very high temperatures.

One example is an electrolytic cell containing molten sodium chloride, NaCl(l), where no water is present. NaCl has a melting point of 801°C, so the cell must be maintained at temperatures greater than this. We can use the OX Z RED method to predict the products of this electrolysis reaction, as shown in Worked example 7.1A.

OX Z RED
an acronym to remember that oxidation is the top half-equation on the electrochemical series, the Z dictates the reactants and products, and reduction is the bottom half-equation

molten
the state of a substance that has been melted to form a liquid

7.1A WORKED EXAMPLE

PREDICTING MOLTEN ELECTROLYSIS REACTIONS

Determine the products formed when molten sodium chloride, NaCl, is electrolysed. Write the balanced overall redox equation for this reaction.

Solution

Think	Do
Step 1: Find the appropriate half-equations on the electrochemical series.	$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(l)$ $E^0 = +1.36$ $Na^+(l) + e^- \rightleftharpoons Na(l)$ $E^0 = -2.71$
Step 2: Identify the chemical species that are reacting.	$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(l)$ $E^0 = +1.36$ $Na^+(l) + e^- \rightleftharpoons Na(l)$ $E^0 = -2.71$
Step 3: Use OX Z RED to help you determine the electrolysis reactions occurring.	$Cl^-(l)$ will undergo oxidation to $Cl_2(g)$ and $Na^+(l)$ will undergo reduction to Na(l).
Step 4: Write the redox half-equations.	Oxidation: $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$ Reduction: $Na^+(l) + e^- \rightarrow Na(l)$
Step 5: Combine the half-equations to write an overall redox equation. Remember to make sure both half-equations have the same number of electrons.	$2Cl^-(l) + 2Na^+(l) \rightarrow Cl_2(g) + 2Na(l)$ The products are chlorine gas and sodium metal. The cell is molten, so the states are all (l).

Study tip

All solid and aqueous states found on the electrochemical series are liquid in molten cells. In half-equations, always replace (aq) and (s) with (l), unless the question directs you to do otherwise.

The positive $Na^+(l)$ ions are attracted to the negative cathode, where they undergo reduction to form sodium metal, Na(l). The negative $Cl^-(l)$ ions are attracted to the positive anode, where they lose an electron to become chlorine gas, $Cl_2(g)$.

All the reactants and products are contained within the same cell, so a porous screen is used to separate the sodium metal from the chlorine gas to prevent them from reacting spontaneously and re-forming sodium chloride, NaCl(l). The pores in the screen allow ions to pass from one side of the cell to the other, but they do not allow the products to pass through.

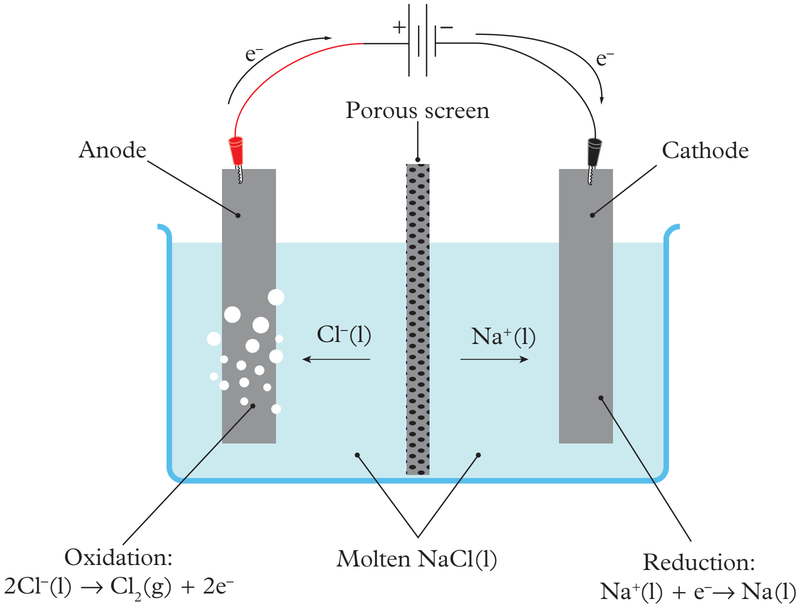


FIGURE 3 The molten NaCl electrolytic cell

Study tip

The products of an electrolytic cell must be kept separate so that they don't react spontaneously. The products of some electrolysis reactions can react violently if they come into direct physical contact with each other.

The voltage that needs to be applied to the electrolytic cell can be determined by calculating the voltage generated by the galvanic cell, as shown in Worked example 7.1B.

7.1B WORKED EXAMPLE

CALCULATING CELL VOLTAGE FOR ELECTROLYTIC CELLS

Determine the voltage that must be applied to a molten sodium chloride electrolytic cell.

Solution

Think	Do
Step 1: Identify the half-reactions occurring in the *electrolytic* cell.	Oxidation: $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$ Reduction: $Na^+(l) + e^- \rightarrow Na(l)$
Step 2: Identify the half-reactions occurring in the *galvanic* cell. This is the opposite of the half-reactions in the electrolytic cell.	Oxidation: $Na(l) \rightarrow Na^+(l) + e^-$ Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(l)$
Step 3: Use the E^0 equation to determine the voltage produced by the galvanic cell. The voltage applied to the electrolytic cell must be *greater* than this value.	$E^0_{\text{cell}} = E^0_{\text{reduction}} - E^0_{\text{oxidation}}$ $= +1.36 - (-2.71)$ $= 4.07\ V$ Therefore, a voltage *greater* than 4.07 V must be applied to the cell.

Study tip

The voltage applied to an electrolytic cell during the recharge must be *greater* than the E^0_{cell} generated in the galvanic cell during discharge. If the galvanic voltage is 2.09 V, it is more accurate to state that a voltage *greater than* 2.09 V is required than that a voltage of 2.10 V is required.

Aqueous electrolysis

It is quite impractical to use the extremely high temperatures required for molten electrolysis, so **aqueous** electrolyte solutions are often used instead. In the electrolysis of aqueous solutions, water can also be a reactant. The water half-equations that are able to undergo electrolysis are:

aqueous
the state of a substance that has been dissolved in water

$$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l) \qquad E^0 = +1.23\,V$$

$$2H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq) + H_2(g) \qquad E^0 = -0.83\,V$$

However, this means that there is now competition between reactants. Let's look at an electrolytic cell containing sodium chloride dissolved in water, NaCl(aq). There are four possible half-equations. The reactants are highlighted:

$$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq) \qquad E^0 = +1.36\,V$$

$$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l) \qquad E^0 = +1.23\,V$$

$$2H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq) + H_2(g) \qquad E^0 = -0.83\,V$$

$$Na^+(aq) + e^- \rightleftharpoons Na(s) \qquad E^0 = -2.71\,V$$

Chloride ions and water can both undergo oxidation, but only the strongest reducing agent will react. On the electrochemical series, water is lower on the right-hand side so it is a stronger reducing agent than chloride ions and will undergo oxidation during electrolysis.

Sodium ions and water can undergo reduction, but only the strongest oxidising agent will react. On the electrochemical series, water is higher on the left-hand side so it is a stronger oxidising agent than sodium and will undergo reduction during electrolysis.

We can fit this into the OX Z RED configuration to find the oxidation and reduction half-equations shown in Figure 4.

The overall reaction equation is:

$$6H_2O(l) \rightarrow O_2(g) + 4OH^-(aq) + 4H^+(aq) + 2H_2(g)$$

$4OH^-(aq)$ and $4H^+(aq)$ form $4H_2O(l)$, so the equation can be further simplified to:

$$2H_2O(l) \rightarrow O_2(g) + 2H_2(g)$$

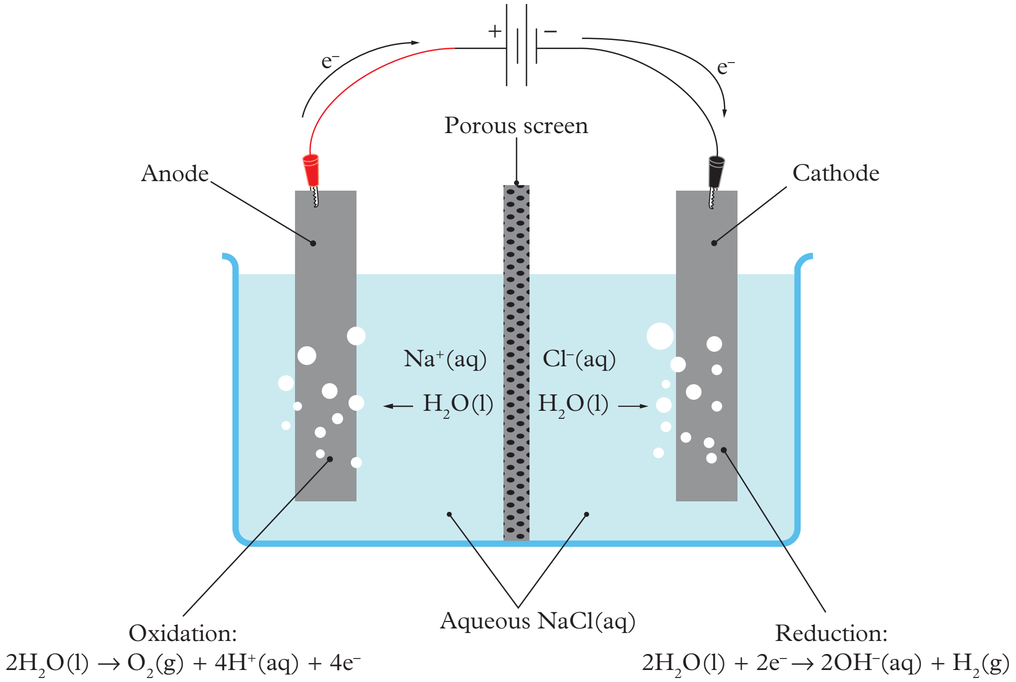


FIGURE 4 The aqueous NaCl electrolytic cell

Reactive electrodes

In molten and aqueous electrolysis, the reactants are supplied via the electrolyte. Reactants for electrolysis can also be the electrodes themselves. One example is electrolysis to purify blister copper, consisting of 98–99.5% copper, contaminated with bubbles of sulfur dioxide and other metals. The impure copper is the anode and pure copper is the cathode. The electrodes are placed in a solution of aqueous acidified copper sulfate. The relevant half-equations are:

$$\begin{aligned}
&O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l) && E^0 = +1.23\,V\\
&Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s) && E^0 = +0.34\,V\\
&Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s) && E^0 = +0.34\,V\\
&SO_4^{\ 2-}(aq) + 2H^+(aq) + 2e^- \rightleftharpoons SO_2(g) + 2H_2O(l) && E^0 = +0.17\,V\\
&2H^+(aq) + 2e^- \rightleftharpoons 2H_2(g) && E^0 = \ \ 0.00\,V\\
&2H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq) + H_2(g) && E^0 = -0.83\,V
\end{aligned}$$

Study tip

Remember to include the half-equations for H_2O if aqueous electrolytes are used, and the half-equations for H^+ if the electrolyte is acidified.

Using the OX Z RED configuration, the strongest oxidising agent is $Cu^{2+}(aq)$ and the strongest reducing agent is Cu(s), so the redox equations are:

Oxidation at the anode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Reduction at the cathode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

Any copper which plates onto the pure copper cathode from the solution will be replaced as the copper in the blister breaks down. Therefore, the concentration of copper ions in solution does not change.

The impurities in the blister copper are also affected by the electrolysis process. Silver and gold are metals that are above copper on the electrochemical series. Therefore, they will remain solid and gather at the base of the container. Any metal that is below copper on the electrochemical series will become an ion and move into the solution.

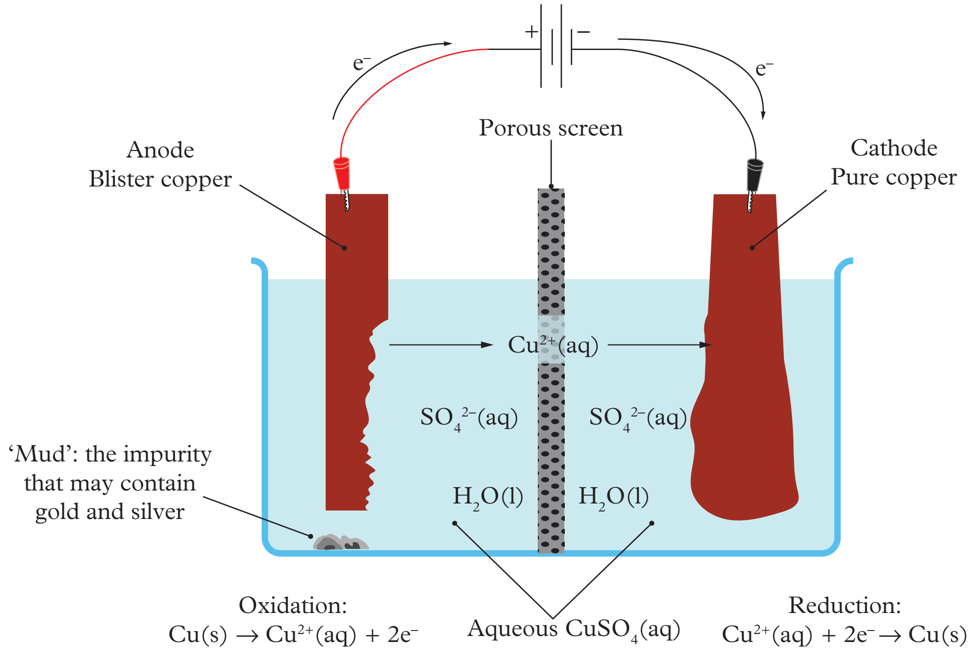


FIGURE 5 The electrolytic cell for purifying blister copper

Limitations of using the electrochemical series

All standard electrode potentials on the electrochemical series are measured at standard conditions (1 M, 100 kPa and 25°C). The oxidising and reducing strength of chemical species are dependent on temperature, concentration and pressure. Operating electrolytic cells under non-standard conditions can result in the swapping of the position of species on the electrochemical series.

For aqueous sodium chloride, the E^0 values of the chloride and water half-equations are very close together:

$$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq) \qquad E^0 = +1.36\,V$$

$$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l) \qquad E^0 = +1.23\,V$$

At non-standard conditions (e.g. less or more than 25°C, 1 atm pressure, 1 M concentration of reactants), the electrolytic cell may produce chlorine gas, rather than oxygen, as the two half-equations swap places. The chloride ion becomes the strongest reducing agent and becomes oxidised in preference to water.

In addition to this, the electrochemical series does not provide information about the rate of the chemical reaction.

7.1 CHECK YOUR LEARNING

Describe and explain

1 Explain why electrolysis reactions are considered non-spontaneous redox reactions.

2 Explain why there are some situations in which molten electrolysis is preferred over aqueous electrolysis.

FIGURE 6 Molten mixtures require an enormous amount of heat to prepare.

3 Using an example, describe the limitations of the electrochemical series for predicting products of electrolysis.

Apply, analyse and compare

4 Compare the movement of electrons in an electrolytic cell with those in a galvanic cell.

5 For the following electrolytic cells:

- **i** identify the strongest oxidising agent and strongest reducing agent
- **ii** write the balanced oxidation and reduction half-equations
- **iii** write the balanced overall equation.

a Molten potassium bromide

b A molten mixture of lithium fluoride and magnesium chloride

c Aqueous magnesium chloride

d Aqueous acidified (in HNO_3) hydrogen peroxide

7.2

Features of commercial electrolytic cells

KEY IDEAS

In this topic, you will learn that:

- commercial electrolytic cells are used to purify metals or plate them onto other metals
- products of electrolysis need to be removed from the cell to prevent them from reacting spontaneously
- the design of an electrolytic cell includes selecting a suitable electrode material, electrolyte and chemical additives.

electrolytic cell
a type of electrochemical cell where electrical energy is converted to chemical energy

7.2 Real-world chemistry
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Electrolytic cells are electrochemical cells that operate in the opposite way to galvanic cells. Instead of converting chemical energy to electrical energy in a spontaneous redox reaction, they convert electrical energy to chemical energy. You learnt a bit about electrolytic cells in Topic 7.1.

Electrolytic cells are used in many industries, to purify metals or plate one onto another, in batteries, beauty treatments (see Real-world chemistry 7.2 in your obook) and the cleaning of water. In this topic, you will explore the features of commercial electrolytic cells and what they are used for.

Electrolytic cells

Commercial electrolytic cells use a form of electrolysis which is designed to purify a substance or extract it from a mixture. The substance, once produced, must be removed from the electrolytic cell so that:

- it does not react spontaneously with other chemical products
- it can be used for other purposes.

The main similarity between electrolytic and galvanic cells is that oxidation occurs at the anode and reduction occurs at the cathode. However, there are several differences, summarised in Table 1.

TABLE 1 Differences between electrolytic and galvanic cells

Electrolytic cell	Galvanic cell
Converts electrical energy to chemical energy	Converts chemical energy to electrical energy
Non-spontaneous reaction	Spontaneous reaction
Uses a single container with a porous divider between the products	Uses separate containers with a salt bridge between them
Keeps products separate so they don't react spontaneously	Keeps reactants separate so they don't react spontaneously
Anode is positive; cathode is negative	Anode is negative; cathode is positive
The positive terminal of the battery is connected to the positive electrode; negative terminal is connected to the negative electrode.	The positive terminal is connected to the negative electrode; negative terminal is connected to the positive electrode.
Cations migrate to the cathode; anions migrate to the anode.	Cations migrate to the cathode; anions migrate to the anode.
Is primarily used to purify metals or plate them onto other metals	Is primarily used to generate useful electrical energy

7.2A
PRACTICAL: CONTROLLED EXPERIMENT

How can you electroplate copper onto a copper cathode? Go to your **Student obook pro**.

7.2B
PRACTICAL: CASE STUDY

Which is best ... molten or aqueous? Go to page **513**.

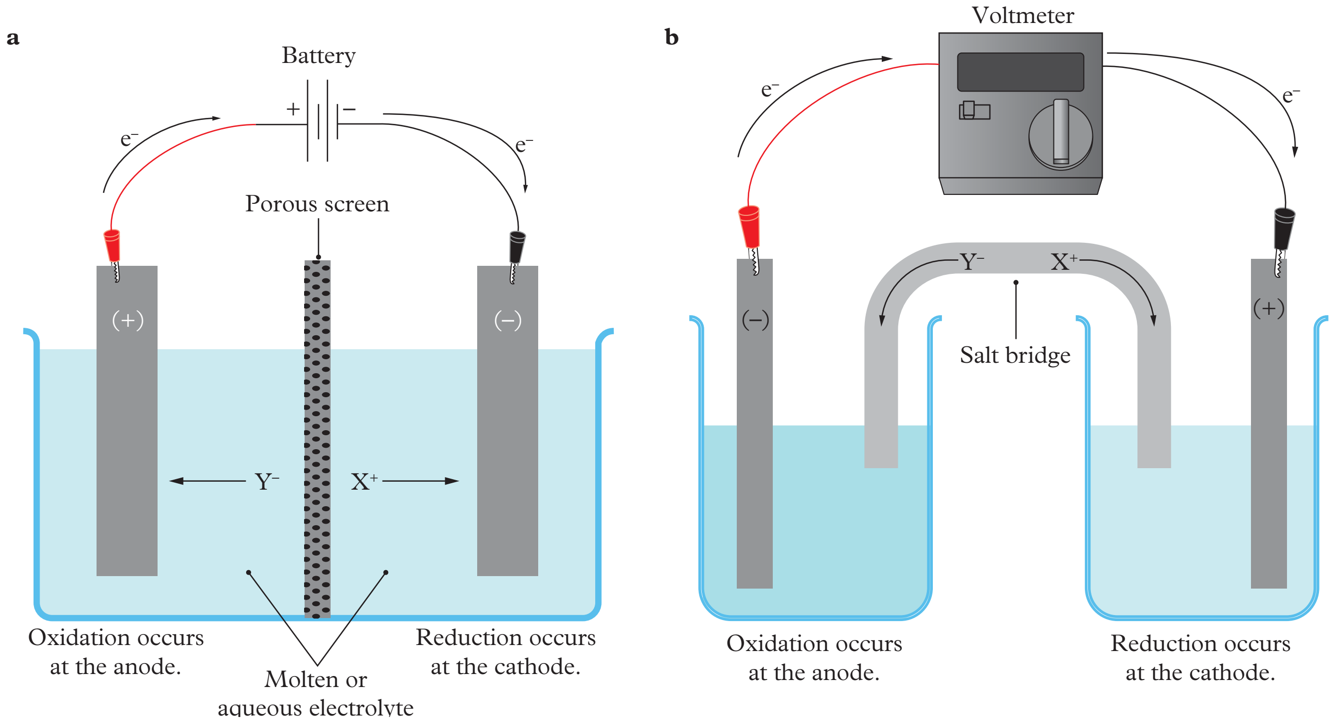


FIGURE 1 **a** Electrolytic and **b** galvanic cells

Designing electrolytic cells

To design electrolytic cells, it is important to consider suitable electrode materials, electrolytes, removal of products as they form, and the use of chemical additives.

Electrodes

Electrodes form a part of the external circuit. They are solid materials which must conduct electrical charge. Electrodes can be reactive or inert.

electroplating
the process in which one metal is coated with another using an electrolysis reaction

Reactive anodes need to be replaced from time to time, since they break down during oxidation in electrolysis. These are often used in a process called **electroplating**, where one metal is covered with another. The ions produced at the anode plate onto the cathode.

When the cathode is susceptible to corrosion, we often coat it in a small layer of a corrosion-resistant metal, which can be expensive. Electroplating is a better option to produce a product that looks the same and is resistant to corrosion. For example, iron is cheap to buy and easy to smelt into various objects. However, it is sensitive to corrosion and can react to form iron oxide (rust). To prevent rusting, silver is electroplated onto iron cutlery and tin is electroplated onto an iron fork (Figure 2). Both metals are unreactive and prevent the rusting of the iron.

In other cells, an inert or unreactive electrode is an advantage. Electrodes made of materials such as graphite or platinum are commonly used in molten cells, where high temperatures are used. For example, aluminium melts at 660°C, whereas NaCl melts at 800°C. You won't be able to use a solid aluminium electrode with a molten NaCl electrolytic cell, because the high temperature will melt the metal electrode. Instead, you would use an inert electrode.

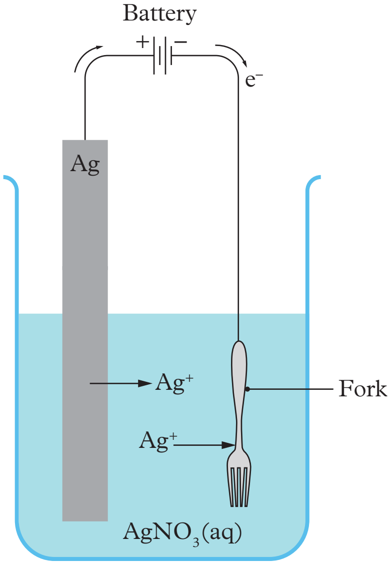


FIGURE 2 Electroplating silver onto an iron fork

Electrolytes

Electrolytes are found in the internal circuit and require free-moving ions to transfer charge between the cathode and anode. This can be done in an aqueous state, where solutes are dissolved in a water solvent, or molten state, where ionic salts are melted.

The strongest oxidising agent will migrate to the cathode to undergo reduction, and the strongest reducing agent will migrate toward the anode to undergo oxidation.

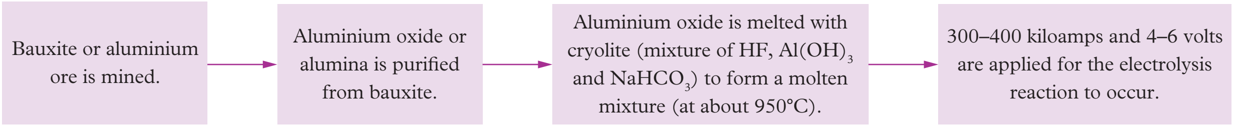
Removal of products

Remember that the reactants in electrolysis do not react spontaneously. Their products, however, will react spontaneously and therefore must be kept separate at each electrode. Usually, this is by using a porous screen.

electropurification the process in which a pure metal is extracted from an impure mixture using an electrolysis reaction

An alternative is to remove products from the cell entirely to avoid this spontaneous reaction. Removal of products is preferred when liquid or gaseous products are formed, because they can be removed more readily. One example is the Hall–Héroult cell for the **electropurification** of aluminium, where the impure aluminium is oxidised to its ionic form (at the anode) and gains electrons to form pure aluminium (at the cathode).

The process is summarised in the flowchart below.



Bauxite or aluminium ore is mined.

Aluminium oxide or alumina is purified from bauxite.

Aluminium oxide is melted with cryolite (mixture of HF, $Al(OH)_3$ and $NaHCO_3$) to form a molten mixture (at about 950°C).

300–400 kiloamps and 4–6 volts are applied for the electrolysis reaction to occur.

FIGURE 3 The Hall–Héroult process

The relevant half-equations are:

$$F_2(g) + 2e^- \rightleftharpoons 2F^-(l)$$

$$CO_2(g) + 4e^- \rightleftharpoons 2O^{2-}(l) + C(s)$$

$$Al^{3+}(l) + 3e^- \rightleftharpoons Al(l)$$

$$Na^+(l) + e^- \rightleftharpoons Na(l)$$

Al^{3+} is the strongest oxidising agent and O^{2-} is the strongest reducing agent. The overall chemical equation for the process is:

$$4Al^{3+}(l) + 6O^{2-}(l) + 3C(s) \rightleftharpoons 3CO_2(g) + 4Al(l)$$

The aluminium and oxide ions can be combined to form an alternative overall equation:

$$2Al_2O_3(l) + 3C(s) \rightleftharpoons 3CO_2(g) + 4Al(l)$$

The molten aluminium metal that forms is denser than the molten cryolite. It sinks to the bottom of the cell and can be removed by a vacuum siphon. Also note that the cathode is made from carbon and carbon is a reactant in the electrolytic process. This makes it a reactive electrode that must be replaced periodically.

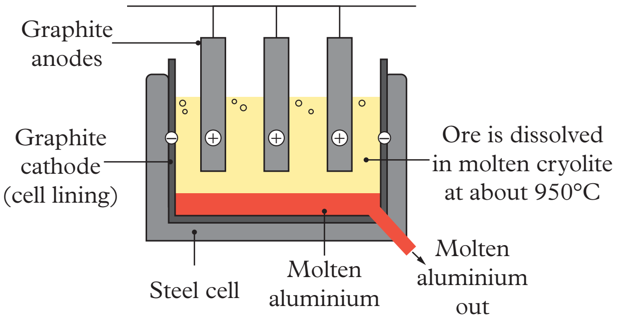


FIGURE 4 The Hall–Héroult electrolytic cell

Chemical additives

In molten electrolysis, the salts in the electrolyte are melted at high temperatures to allow the ions to move freely. The advantage of this is that it avoids using water as a reactant, as water can compete with the other oxidising and reducing agents. The disadvantage is that due to the high temperatures, a lot of energy is needed to operate the cell.

Chemical additives can be introduced to lower the melting point of the electrolyte in a molten electrolysis cell. When you select an additive, it is essential to ensure that it will not compete or react at either electrode instead of the desired chemical reaction. By lowering the melting point of the electrolyte, less energy needs to be applied to the cell.

Have a go at designing an electrolytic cell that produces zinc metal in Challenge 7.2.

chemical additive
a chemical that can be introduced into a system to improve a reaction

7.2 Challenge
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7.2 CHECK YOUR LEARNING

Describe and explain

1 Describe the structural differences between a galvanic cell and an electrolytic cell.

2 Describe one situation in which you would use a reactive electrode over an inert one.

Apply, analyse and compare

3 Explain how impurities such as H_2O and Fe^{2+} would affect the chemical equations occurring in the Hall–Héroult cell. Refer to the strongest oxidising and reducing agents in your response.

4 Explain why it is not possible to conduct the Hall–Héroult cell reaction in an aqueous solution. Use balanced chemical equations and the terms *strongest oxidising agent* and *strongest reducing agent* in your response.

5 Construct an electrolytic cell for the following chemicals. In your diagram, label:

- the anode and the cathode
- electrode material
- electrode polarity
- oxidation and reduction half-equations
- the strongest oxidising agent and the strongest reducing agent
- the overall equation
- the movement of ions in the cell
- the movement of electrons
- the voltage that must be applied to the cell.

a Molten aluminium fluoride

b A molten mixture of zinc bromide, aluminium chloride and nickel fluoride

c Aqueous cobalt iodide

d An aqueous mixture of sodium chloride, magnesium iodide and sulfuric acid

Design and discuss

6 Discuss the environmental impacts of the Hall–Héroult cell. Refer to the green chemistry principles in your response.

7 A scientist needs to use chlorine gas in a reaction and decides to use electrolysis to generate the gas. Design a method to produce chlorine gas under standard conditions. Your answer should include the chemicals and half-equations used, electrode materials, the applied cell voltage and reasons for your choices.

8 An experiment was set up to plate silver metal onto a zinc fork. First, the fork and the silver metal electrode were cleaned with emery paper to remove any products of oxidation. A 1.00 M solution of silver nitrate was placed in a beaker. The positive terminal of the battery was connected to the fork and the negative terminal to the silver electrode. The battery was turned to 4 V and run for 10 minutes.

a Predict the outcome of the experiment based on the method.

b Evaluate the method and discuss anything that may affect the ability to plate silver onto the fork.

c Design an improved method for this experiment.

7.3

Features of rechargeable secondary cells

KEY IDEAS

In this topic, you will learn that:

- secondary cells (rechargeable batteries) can discharge as a galvanic cell and recharge as an electrolytic cell
- a voltage greater than the voltage generated by the cell during discharge must be supplied to the cell to reverse the redox reaction.

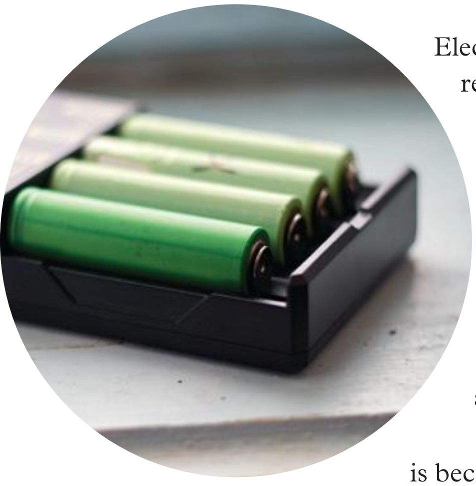

FIGURE 1 Rechargeable batteries

Electrolysis reactions are essential to modern society because they are the basis of rechargeable batteries, which power devices such as cars, laptops and phones. In this topic, you will learn about the features of an electrochemical secondary cell that allow it to be rechargeable.

Secondary cells – rechargeable batteries

Unlike primary cells, which can no longer produce electricity once the reactants have been consumed, **secondary cells** can be recharged. By applying a voltage back into the cell, the redox reactions are reversed, and the battery is returned to almost its original state.

Over time, rechargeable batteries degrade and can store less and less charge. This is because recharging the battery does not convert all of the chemicals to their original state. Skill drill 7.3 gives you a chance to practise your key science skills by designing an investigation to study battery life.

secondary cell
an electrochemical cell that can operate in both discharge (as a galvanic cell) and recharge (as an electrolytic cell)

7.3 Skill drill
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Galvanic is discharge and electrolytic is recharge

The terms *discharge* and *recharge* are frequently used to determine whether a secondary cell is converting chemical to electrical energy or electrical to chemical energy. A cell in:

- **discharge** converts chemical energy to electrical energy
- **recharge** converts electrical energy to chemical energy.

In the discharge mode, the cell operates as a galvanic cell. This reaction is spontaneous, producing products that cannot react spontaneously with each other. If the electrical energy applied to the cell in electrolysis is greater than the energy generated during discharge, the reaction can be reversed. This is the recharge mode, where the cell operates as an electrolytic cell.

Discharge mode
Galvanic cell

$$\text{Reactants} \xrightarrow{\text{chemical} \rightarrow \text{electrical}} \text{Products}$$

Recharge mode
Electrolytic cell

$$\text{Products} \xleftarrow{\text{electrical} \rightarrow \text{chemical}} \text{Reactants}$$

FIGURE 2 The products of a spontaneous reaction form the reactants of the electrolytic reaction.

discharge
conversion of chemical to electrical energy in a redox reaction

recharge
conversion of electrical to chemical energy in a redox reaction

Features and conditions required for cell reactions to be reversed

A secondary cell can be connected to an appliance or a load (a device that consumes electrical energy). When connected to a load, the cell will function as a galvanic cell to discharge electrical energy. When connected to a power source, the cell will function as an electrolytic cell to recharge chemical energy.

Unlike a primary cell, which can convert reactants to products, a secondary cell must allow the products generated to be converted back into reactants during recharge, so that they can generate electrical energy once again. This relies on the:

- presence of a separator
- voltage applied
- reversibility of electrode polarity.

Study tip

The products of electrolysis must be separated, or they will react spontaneously, producing large amounts of heat energy.

Study tip

Discharge is spontaneous and follows the RED SOX method. Recharge is non-spontaneous and follows the OX Z RED method.

Separator

The reactants of recharge (which are the products of discharge) must remain in contact with the electrode. If they are not in contact, then the oxidising agent cannot gain electrons from the cathode to undergo reduction.

However, it is also important for the products of the discharge reaction to be separate from each other. If products come into contact, they will react spontaneously and generate large amounts of heat (instead of electrical) energy, potentially causing harm to the battery and the person operating it.

To prevent this from happening, a separator is used between the cathode and anode. Usually, this is a **porous membrane** made of polymer material. The holes in the membrane only allow positive ions in the electrolyte to move between the electrodes. The negative ions are unable to pass through and remain in each compartment. This **semipermeable** membrane prevents the products from interacting (Figure 3).

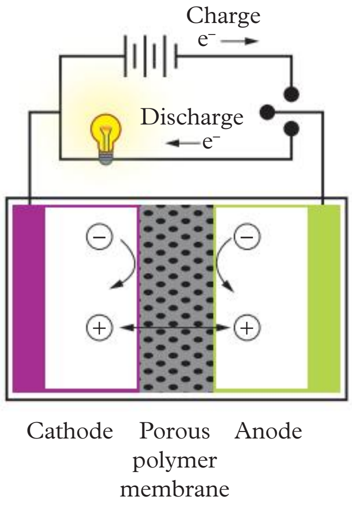

FIGURE 3 A porous membrane separates the products of electrolysis while allowing ions to move in the electrolyte.

porous membrane a barrier that contains holes (pores) for substances to pass through

semipermeable allows some, but not all, substances to pass through

Applied voltage

To recharge a secondary cell, a voltage greater than the voltage produced in discharge must be applied by the battery. If the cell produces 4 V in discharge, it requires a voltage greater than (>) 4 V to be recharged.

Electrode polarity

The electrode polarity must be reversible. When a galvanic cell discharges, electrons move from the negative anode to the positive cathode. This is because negatively charged electrons are produced at the anode. When an electrolytic cell recharges, electrons move from the positive anode to the negative cathode. This is because the battery pulls electrons from the anode and pushes them towards the cathode.

Try Challenge 7.3 to test your understanding of galvanic and electrolytic cell reactions.

7.3 Challenge
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Car batteries – the lead–acid accumulator

One example of a rechargeable battery is the car battery. This is also called a lead–acid accumulator because of its lead electrodes and sulfuric acid (H_2SO_4) electrolyte. The two half-equations occurring in this type of secondary cell are:

$$PbO_2(s) + 4H^+(aq) + SO_4^{\ 2-}(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l) \qquad E^0 = +1.69\,V$$

$$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{\ 2-}(aq) \qquad E^0 = -0.36\,V$$

In discharge, we use RED SOX to determine the reactions occurring.

Anode (–): $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$

Cathode (+): $PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$

Overall reaction: $PbO_2(s) + Pb(s) + 2SO_4^{2-}(aq) + 4H^+(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

The four hydrogen ions ($4H^+$) and two sulfate ions ($2SO_4^{2-}$) on the reactant side can be combined to form two sulfuric acid molecules ($2H_2SO_4$). Therefore, the overall reaction equation can be simplified to:

$$PbO_2(s) + Pb(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$$

When fully charged, the battery has a high level of chemical potential energy. The electrodes are made of lead(II) oxide and lead, and there is an aqueous solution of sulfuric acid as an electrolyte. As the battery discharges, sulfate ions are removed from the electrolyte, decreasing the concentration of the acid, and lead sulfate forms on the electrodes, coating them and stopping the spontaneous reaction (Figure 4).

Study tip

Remember that, unlike galvanic and fuel cells which are only connected to a load, electrolytic cells must also be connected to a power source.

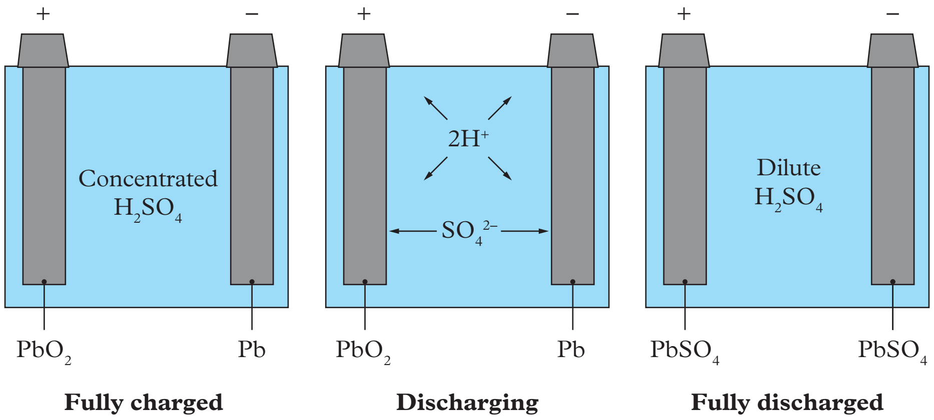

FIGURE 4 The electrodes and electrolyte in a lead–acid battery

Each lead–acid cell can produce a voltage of 2.05 V. When multiple cells are connected in series, the voltage of the battery is the sum of the cells. Six lead–acid cells are connected to produce a 12 V car battery (Figure 5).

FIGURE 5 Six 2 V lead–acid cells are arranged in series to produce a car battery of 12 V.

To recharge the battery, the products of the discharge reaction must remain in contact with the electrode to act as the reactants of the recharge reaction. In addition, a voltage greater than 12 V must be used. The reactions are reversed and follow the OX Z RED method:

Anode (+): $PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^-$

Cathode (–): $PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$

Overall reaction: $2PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + Pb(s) + 2H_2SO_4(aq)$

Lithium–ion batteries

In a lithium–ion battery, lithium ions move between the electrodes.

As the cell discharges, the lithium ions move from the graphite anode, which contains lithium **intercalated** between the graphite sheets (LiC_6), to the metal oxide cathode, which the ions dope (or cover) to form $LiCoO_2$.

intercalated
inserted between layers

Anode (–): $LiC_6(s) \rightarrow 6C(s) + Li^+(aq) + e^-$

Cathode (+): $CoO_2(s) + Li^+(aq) + e^- \rightarrow LiCoO_2(s)$

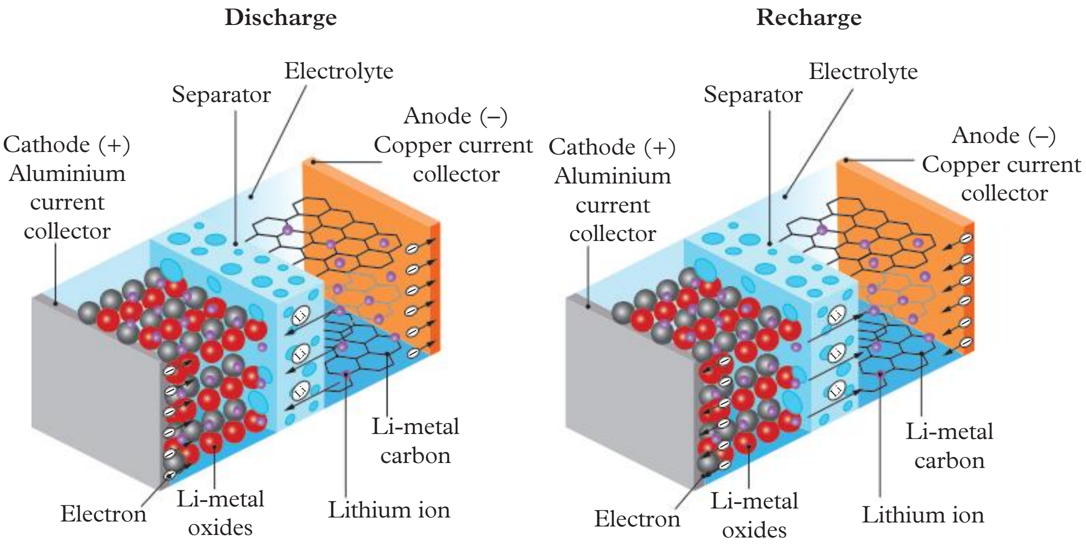

FIGURE 6 The lithium–ion battery

When the battery is recharged, the reactions reverse and the lithium at the metal oxide electrode moves back to the graphite electrode.

$$\text{Anode (+): } 6C(s) + Li^+(aq) + e^- \rightarrow LiC_6(s)$$

$$\text{Cathode (–): } LiCoO_2(s) \rightarrow CoO_2(s) + Li^+(aq) + e^-$$

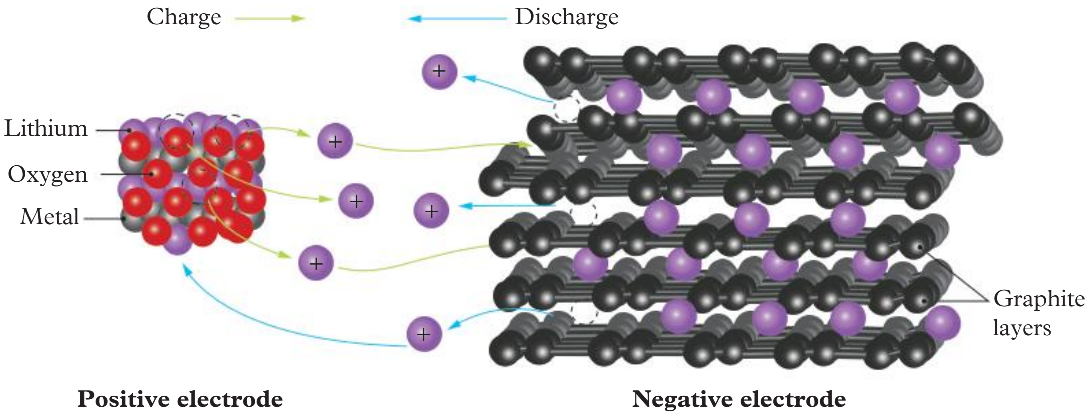

FIGURE 7 The electrodes of a lithium–ion battery, doped with lithium ions (purple)

A typical lithium–ion battery can produce about 4 V and can undergo up to 1200 recharge–discharge cycles before it dies. These types of batteries are some of the most efficient and are used in portable devices, power tools and even electric cars. Use Worked example 7.3 to see how you can use your understanding of secondary cells to answer questions about batteries.

Table 1 on the next page summarises the differences between discharge and recharge in secondary cells.

7.3 Worked example
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7.3 Worked example
Video demonstration

TABLE 1 Discharge and recharge in secondary cells

Mode	Discharge	Recharge
Type of cell	Galvanic	Electrolytic
Type of reaction	Spontaneous	Non-spontaneous
Energy transformation	Chemical to electrical	Electrical to chemical
Considerations for reversibility of the reactions	Products must remain in contact with the electrodes so that the reaction can be reversed.	Products must be separated from one another, or they will react spontaneously, producing heat energy.
Anode (oxidation)	Negative (oxidation produces electrons)	Positive
Cathode (reduction)	Positive	Negative (the power source pushes electrons to the electrode)
Predicting products	RED SOX method	OX Z RED method
Electron transfer	Anode to load to cathode	Anode to power source to cathode
Voltage	Produces electrical energy in discharge. $E^0_{\text{cell}} = E^0_{\text{reduction}} - E^0_{\text{oxidation}}$	Electrical energy supplied must be greater than the E^0_{cell} produced in discharge.

7.3 CHECK YOUR LEARNING

Describe and explain

1 Describe the difference between a primary and a secondary cell.

2 Explain why it is necessary to keep the products of recharge in contact with the electrodes. Predict what would happen if these products fell away from the electrodes.

3 Describe the components of a secondary cell that are required for that the cell to recharge.

Apply, analyse and compare

4 The following half-equations occur in a nickel–cadmium alkaline battery in discharge:

$\text{Cd(s)} + 2\text{OH}^-\text{(aq)} \rightarrow \text{Cd(OH)}_2\text{(s)} + 2\text{e}^- \quad E^0 = +0.81\ \text{V}$

$\text{NiO(OH)(s)} + \text{H}_2\text{O(l)} + \text{e}^- \rightarrow \text{Ni(OH)}_2\text{(s)} + \text{OH}^-\text{(aq)} \quad E^0 = -0.49\ \text{V}$

- **a** Identify the species that undergo oxidation and reduction. Determine which reaction occurs at the anode and which reaction occurs at the cathode.
- **b** Construct the overall chemical equation.
- **c** Construct the half-equations occurring at the anode and the cathode in recharge and identify them as oxidation or reduction.
- **d** Calculate the voltage required to recharge the nickel–cadmium battery.

5 In a vanadium redox flow battery, the following half-equations are used to produce electricity in discharge.

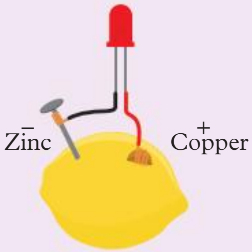
$\text{VO}_2^{\ +}\text{(aq)} + 2\text{H}^+\text{(aq)} + \text{e}^- \rightarrow \text{VO}^{2+}\text{(aq)} + \text{H}_2\text{O(l)} \quad E^0 = +0.99\ \text{V}$

$\text{V}^{3+}\text{(aq)} + \text{e}^- \rightarrow \text{V}^{2+}\text{(aq)} \quad E^0 = -0.26\ \text{V}$

- **a** Construct the half-equation occurring at the anode and the cathode in recharge.
- **b** Construct the overall electrolytic reaction equation.
- **c** Calculate the voltage required to recharge the cell.
- **d** Identify a suitable electrolyte for the cell.

Design and discuss

6 Investigate the lemon battery.



- **a** Determine the half-equations and overall chemical equation occurring in the lemon battery.
- **b** Determine what voltage may be gained from a lemon battery. Identify the types of appliances or devices it could power for.
- **c** Discuss what happens to the battery over time. Compare this to commercial batteries.
- **d** Discuss the term *battery* in reference to the lemon battery.

7.4

Producing 'green' hydrogen

KEY IDEAS

In this topic, you will learn that:

- hydrogen is traditionally produced by steam methane reforming
- 'green' hydrogen can be produced by electrolysis.

In Chapter 4, you learnt that fuel cells are used to provide energy in society. Many of today's fuel cells use hydrocarbon fuels that have been extracted from fossil fuels. Not only is this problematic because fossil fuels are a finite resource, but the burning of hydrocarbons to produce energy also generates carbon dioxide, which contributes to greenhouse gas emissions.

Hydrogen gas is the best alternative to the use of fuels extracted from fossil fuels. The relative amounts of energy produced by hydrogen gas compared to fossil fuels are shown in Table 1.

TABLE 1 Energy produced by fuels

Fuel	Energy produced ($kJ\,g^{-1}$)
Hydrogen gas	141
Methane	55.6
Natural gas	54.0

Despite being the most abundant element in the universe, it occupies only 0.000 05% of the atmosphere in its molecular form, H_2. Every year, approximately 500 billion cubic metres (1 cubic metre = 1000 L) of hydrogen gas is produced. Because more than 95% of the hydrogen in the world is produced from non-renewable sources, significant importance is placed on the development of green methods to produce hydrogen gas on an industrial scale.

Steam methane reforming

steam methane reforming
a process in which methane is heated with steam and a catalyst to produce carbon monoxide and hydrogen gas; also called steam methane cracking (SMC)

'green' hydrogen
hydrogen produced by chemical techniques that are efficient, require minimal energy, produce no waste and are carbon-neutral

electrosynthesis
a process in which electrolysis is used to form new products

The process of **steam methane reforming** (SMR) involves heating methane with steam (vapourised water) and a catalyst to produce carbon monoxide and hydrogen gas:

$$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)$$

Carbon monoxide, a highly toxic gas, is reacted with water to produce more hydrogen and carbon dioxide:

$$CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$$

This process is not sustainable because every atom of carbon present in the initial methane reactant is converted to carbon dioxide. The extraction of methane from natural gas also releases carbon dioxide, further contributing to greenhouse gas levels.

Hydrogen gas produced by SMR is impure due to contaminants present within fossil fuels, CO and CO_2. However, it is cheaper and faster, and there are already established production plants.

Electrosynthesis of 'green' hydrogen

Recent developments in electrolysis focus on greener methods of producing hydrogen fuel. The production of **'green' hydrogen** uses **electrosynthesis** to form products that may otherwise be difficult to form from other chemical reactions, or that would be impractical to produce due to the many reaction steps required.

Alkaline electrolysis – water splitting

water splitting
an alkaline electrolysis reaction whereby water is split into oxygen and hydrogen gases

In an electrolytic process, the two electrodes are immersed in an alkaline aqueous electrolyte and electrical energy is applied to produce oxygen and hydrogen gases. This is called alkaline electrolysis or **water splitting**, because it splits the water into oxygen and hydrogen gases.

At the negative cathode, the water molecules are reduced to form hydrogen gas and hydroxide ions:

$$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$$

At the positive anode, hydroxide ions are oxidised into oxygen gas and water:

$$4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$$

The process requires a constant supply of water, an alkaline electrolyte (preferably NaOH or KOH) and a voltage greater than 1.23 V. It is greener than SMR because it does not release carbon into the atmosphere. However, the electrical energy must be sourced from a power plant. The overall equation for this process is:

$$2H_2O(l) \rightarrow O_2(g) + 2H_2(g)$$

This is the reverse of the process that occurs in alkaline fuel cells, which you learnt about in Chapter 4. Can you draw a diagram to show the electrolysis of water under alkaline conditions?

Polymer electrolyte membrane electrolysis

PEM electrolysis
water electrolysis with an acid electrolyte used to produce hydrogen gas using a polymer electrolyte membrane

Polymer electrolyte membrane (PEM) electrolysis is an innovation in 'green' hydrogen synthesis. It produces hydrogen with a purity of approximately 99.99%. The cell uses a solid polymer membrane electrolyte that:

- conducts hydrogen ions from the anode to the cathode
- separates oxygen and hydrogen gases
- provides electrical insulation to the electrodes.

Although the membrane is acidic, there are no solutions in this cell as all components are solid. The electrodes are also often embedded with catalysts to accelerate the reaction.

Water is pumped into the cell at the anode. The oxidation half-reaction that occurs at the anode is referred to as the oxygen evolution reaction (OER).

The water splits into oxygen gas (which exits the cell) and hydrogen ions (which move through the membrane to the cathode):

$$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$$

At the cathode, the hydrogen ions are reduced to form hydrogen gas. This is also called the hydrogen evolution reaction (HER).

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

The energy required for the electrolysis reaction (>1.23 V) is generated by photovoltaic (solar) or wind energy. This reduces the cell's carbon footprint because no carbon dioxide is produced and released into the atmosphere as a by-product. In homes that use solar power, excess energy generated by photovoltaic cells can be used to produce hydrogen gas, which can then be stored in a gas tank for later use. However, due to the use of expensive catalysts (e.g. iridium, platinum, titanium, gold), the cell is costly.

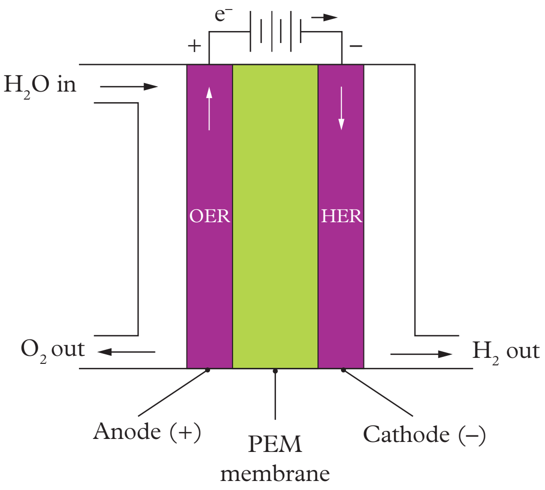


FIGURE 1 PEM electrolytic cell

Artificial photosynthesis

Photosynthesis is an electrochemical process that occurs in the chlorophyll of leaves in plants. Water from the soil and carbon dioxide from the atmosphere are converted into glucose using sunlight:

$$CO_2(g) + H_2O(l) \rightarrow C_6H_{12}O_6(aq) + O_2(g) \qquad \Delta H = +2816\,kJ\,mol^{-1}$$

Photosynthesis occurs in two phases. Phase 1 is the light-dependent reaction, also called water splitting or water oxidation. Phase 2 is the light-independent reaction, where carbon dioxide is converted to glucose. Scientists recognised that the first phase produces hydrogen gas, so they have worked to take advantage of this reaction as a greener way of producing hydrogen. This process is called **artificial photosynthesis**.

Water oxidation and the proton reduction catalyst system

Artificial photosynthesis occurs in **photoelectrochemical cells** (PECs). This is also called a water oxidation and proton reduction catalyst system because of the reactions that occur.

The anode must be coated in a catalyst, such as titanium oxide (TiO_2), which can absorb **photons** (hv) of sunlight. The photons cause the catalyst to release an electron and leave a positive electron hole behind (h^+), as shown in Figure 2. This is called the photovoltaic effect.

$$hv \rightarrow h^+ + e^-$$

The electrons travel to the cathode and the water molecules interact with the holes that are left behind at the anode to form oxygen and protons (H^+):

$$4h^+ + 2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$$

At the cathode, protons are reduced to form hydrogen gas:

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

artificial photosynthesis
the production of hydrogen gas by water splitting and proton catalyst reduction using light energy

photoelectrochemical cells
electrochemical cells that use light as a reactant to produce hydrogen gas

photon
an uncharged particle that is the smallest possible amount of electromagnetic or light energy

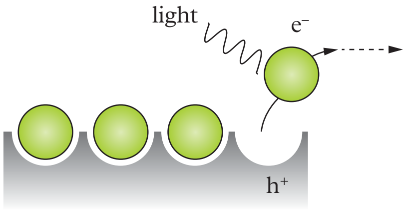
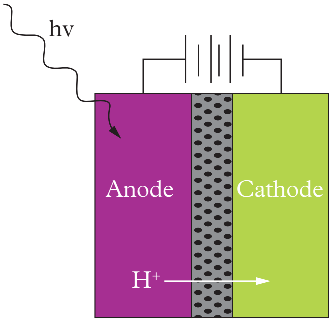


FIGURE 2 The photovoltaic effect



$$4h^+ + 2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$$
$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

FIGURE 3 A photoelectrochemical cell

Table 2 summarises the four methods to produce hydrogen gas that we've looked at, and compares their energy input, efficiency of energy transformation, and estimated cost.

TABLE 2 Methods to produce hydrogen gas

Method	Energy input	Reaction efficiency (%)	Cost (per kg of H_2)
Steam methane reforming	Electrical energy from a power plant	74–85	\$2.30
Alkaline electrolysis	Electrical energy from a power plant	60–80	\$10.30
Polymer electrolyte membrane electrolysis	Photovoltaic or wind energy	80–90	\$5–10
Artificial photosynthesis	Photovoltaic energy	10–30	Estimated \$10

Although there are many advantages to electrolysis, the disadvantages are also widely acknowledged. The processes are expensive and require large amounts of electrical or thermal energy. Processes are constantly being improved and optimised to reduce waste chemicals, reduce energy consumption and reduce cost. The 2020 goal for hydrogen uses the slogan ‘1 1 1’. This means they are aiming for hydrogen to be produced at $1 per 1 kilogram in 1 decade (by 2030).

7.4 SKILL DRILL

Exploring the sustainability of energy production

Key science skill: Analyse, evaluate and communicate scientific ideas

A common misconception is that secondary cells provide cleaner energy because they are rechargeable, and they do not produce any waste or generate harmful atmospheric gases such as carbon dioxide. However, when they are recharged, secondary cells use mains power. In Australia, this is mostly generated by burning coal, which produces harmful greenhouse gases that contribute to global warming and climate change.

Practise your skills

1 Compare the sustainability of primary and secondary cells, in terms of linear and circular economies.
2 Discuss how ‘green’ hydrogen could be a better alternative to meeting society’s energy needs.
3 Identify the green chemistry principles that relate to the production of ‘green’ hydrogen.

Need help applying sustainability perspectives? See Topic 1.11 (page 32).

7.4 CHECK YOUR LEARNING

Describe and explain

1 Explain why it is necessary to produce hydrogen gas.
2 Explain what ‘green’ hydrogen is.

Apply, analyse and compare

3 Using a calculation, show that combustion of 1 g of hydrogen produces more energy than 1 g of methane or natural gas.
4 Compare alkaline electrolysis and artificial photosynthesis.
5 Construct an electrolytic cell for alkaline electrolysis. You can use Figure 1 as a guide. Make sure to include:
   - the anode and the cathode
   - electrode materials
   - electrode polarity
   - oxidation and reduction half-equations
   - the strongest oxidising agent and strongest reducing agent
   - the overall equation
   - the movement of ions in the cell
   - the movement of electrons
   - the voltage that must be applied to the cell.

Design and discuss

6 Discuss ‘green’ hydrogen synthesis. Explain why electrolytic cells are considered green, but steam methane reforming is not. Refer to sustainability perspectives in your response.
7 Discuss why steam methane reforming is still used as the predominant method of hydrogen production.

7.5

Faraday’s Laws in electrolysis

KEY IDEAS

In this topic, you will learn that:

- Faraday’s Laws can be used to predict the quantity of reactant and product or the amount of current or time required to consume a certain quantity of reactant or produce a certain quantity of product.

FIGURE 1 Michael Faraday developed Faraday’s Laws based on his electrolysis experiments.

Faraday’s Laws of electrolysis were first developed using electrolytic cells. These laws have been covered in detail in Topic 4.6. In the same way that you used Faraday’s Laws to predict the quantities of reactants used and products formed in spontaneous redox reactions, you can also apply them to the reactants and products of electrolysis.

Applying Faraday’s Laws to electrolysis

In electrolysis, Faraday’s Laws and stoichiometry are useful for many types of calculations, including determining:

- the mass of product formed
- the quantity of reactant electrolysed or remaining after a reaction
- the amount of time or current required for a reaction to produce a certain mass of metal.

Below is a quick recap of Faraday’s Laws, but you are encouraged to go back to Topic 4.6 to revise this if you have forgotten.

Faraday’s first law states that the charge applied to the cell is proportional to the mass of metal plated onto the cathode:

$$Q \propto m$$

where Q is the charge and m is the mass of the metal deposited onto the cathode.

Faraday’s second law states that, to produce 1 mole of a metal, 1, 2, 3 or another whole number of moles of electrons must be consumed:

$$Q = n(e^-) \times F \quad \text{or} \quad Q = n(e^-) \times 96\,500$$

where Q is the electric charge, $n(e^-)$ is the number of moles of electrons and F is the Faraday constant (the charge on one mole of electrons).

One more equation that will be useful is the electric charge equation:

$$Q = I \times t$$

where Q is the electric charge in coulombs, I is the current in amps and t is the time in seconds. We can also look at a few worked examples on the next page to see how to use Faraday’s Laws in electrolysis reactions. You can use Skill drill 7.5 to practise your electrolysis data analysis skills.

Study tip

Remember to use a ratio statement when connecting the number of moles of electrons to the number of moles of the substance at the anode or cathode.

7.5 Skill drill

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Study tip

If the half-equation is not given to you in the question, you *must* provide one in your answer.

7.5A WORKED EXAMPLE

CALCULATING THE MASS OF AN ELECTRODE AFTER ELECTROPLATING

An aqueous copper sulfate electroplating cell was made using 1.0 M $CuSO_4$ and two 1.50 g copper electrodes. A voltage of 8.0 V is applied, creating a current of 5.00 A for 5.00 minutes. Calculate the mass of the cathode after the five-minute time period.

Solution

Think	Do
Step 1: Calculate the charge applied to the electroplating cell.	$Q = I \times t$ $= 5.00 \times (5.00 \times 60)$ $= 1500\,C$
Step 2: Calculate the amount (in mol) of electrons produced.	$n(e^-) = \frac{Q}{F}$ $= \frac{1500}{96\,500}$ $= 0.0155\,mol$
Step 3: An aqueous cell will contain H_2O as a potential reactant. Use the electrochemical series to confirm the half-reactions occurring. Never assume that a reaction will occur.	$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$ $E^0 = +1.23\,V$ $Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$ $E^0 = +0.34\,V$ $2H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq) + H_2(g)$ $E^0 = -0.83\,V$ $Cu^{2+}(aq)$ is the strongest oxidising agent and Cu(s) is the strongest reducing agent, so water will not react.
Step 4: Write the half-equation for electroplating. If this is not given to you in the question, you *must* write one.	Solid copper is being plated onto the electrode, so the reaction is: $Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$
Step 5: Use a ratio statement to determine the amount (in mol) of copper produced.	$n(Cu) = n(e^-) \times \frac{1}{2}$ $= 0.0155 \times \frac{1}{2}$ $= 7.77 \times 10^{-3}\,mol$
Step 6: Calculate the mass of copper plated onto the electrode.	$m(Cu) = n \times M$ $= 7.77 \times 10^{-3} \times 63.5$ $= 0.49\,g$
Step 7: Add the mass of the copper to the mass of the cathode.	$1.5 + 0.49 = 1.99\,g$ (3 sig fig)

Faraday's Laws can also be used to determine the charge on a metal ion. This is handy when you have a metal that can have multiple oxidation states. See how to do this in Worked example 7.5D.

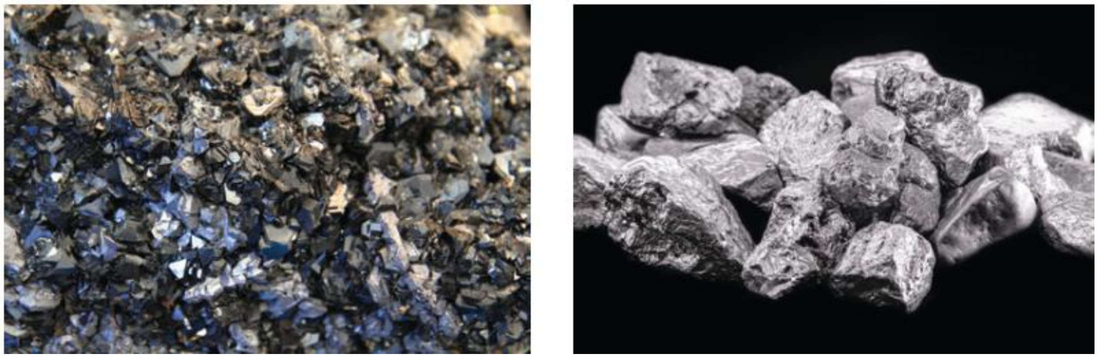

FIGURE 2 Lead and cobalt are metals that can form ions with different charges.

7.5B Worked example
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7.5B Worked example
Video demonstration

7.5C Worked example
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7.5C Worked example
Video demonstration

7.5D Worked example
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7.5D Worked example
Video demonstration

7.5 CHALLENGE

Calculating the thickness of plating on an electrode

A tin mug is given as an award for the Top Scoring Student in VCE Chemistry. Before it is presented, it must be plated with silver. The mug has a surface area of 280 cm^2, and the density of silver is 10.49 g cm^{-3}. Assuming that there is only 1.00 hour to plate the silver onto the mug and that a current of 5.0 A can be generated, calculate the thickness of silver that can be plated onto the mug.

$$\text{thickness of the metal (cm)} = \frac{\text{volume (cm}^3\text{)}}{\text{surface area (cm}^2\text{)}}$$

1 Calculate the mass of the silver that is plated.

2 Calculate the volume of silver based on its mass and density.

3 Calculate the thickness of the silver using the equation provided.

7.5 CHECK YOUR LEARNING

Describe and explain

1 Describe what kind of information is required if you want to calculate the mass of reactant produced by an electrolytic cell.

2 Explain why you can use Faraday's Laws to calculate the quantities of reactants and products in both galvanic and electrolytic cells.

Apply, analyse and compare

3 An aqueous nickel chloride electroplating cell was made using 1.0 M $NiCl_2$. Both electrodes are made of silver. A voltage of 12.0 V is applied, creating a current of 2.50 A for 23.50 minutes. Calculate the mass of product generated at the cathode after the 23.50 minutes time period.

4 In an electrolytic cell, 150.0 mL of a 2.0 M lead(II) nitrate ($Pb(NO_3)_2$) solution is electrolysed using a current of 3.60 A for 12.00 minutes. Both electrodes are made of graphite.

   a Calculate the mass of product deposited at the cathode.

   b Calculate the new concentration of the lead(II) nitrate solution.

5 An aqueous magnesium fluoride electrolytic cell was made using 1.00 M MgF_2. Both electrodes are made of graphite. A voltage of 7.50 V is applied, creating a current of 4.5 A for 15.50 minutes. Calculate the mass of product generated at the cathode after 15.50 minutes if the cell is run at SLC.

6 To plate 2.50 g of zinc onto the cathode of a cell, a current of 18.0 A is applied. Calculate how long (in hours) the cell has to operate to deposit this mass onto the electrode.

7 To determine the charge on a metal ion, a current of 17.19 A is run for 10 minutes to plate a mass of 1.36 g of vanadium at the cathode of a cell. Determine the charge on the vanadium ion.

Design and discuss

8 A student is given a sample of vanadium nitrate salt in a laboratory. They are given the task of determining the formula of the chemical compound, but know that vanadium can be a V^{5+}, V^{4+}, V^{3+} or V^{2+} ion.

   a Design an experiment that the student can use to determine the charge on the vanadium ion.

   b Assuming that the ion being analysed is the V^{3+} ion, based on the current and voltage that you used for part **a**, calculate the mass of vanadium that you would expect to measure.

CHAPTER 7 Review

Chapter summary

7.1

- Electrolysis involves passing an electrical current through a galvanic cell to reverse the spontaneous chemical reactions.
- The polarity of electrodes is reversed compared to galvanic cells as electrons are drawn away from the anode, where oxidation occurs, making it positive. Electrons are then pushed onto the cathode, making it negative. The cathode becomes the site of reduction, because electrons are consumed at this electrode.
- The electrochemical series can be used to predict the products of electrolysis reactions, as long as the state (molten liquid or in aqueous solution) and the electrode materials are known.

7.2

- Commercial electrolytic cells apply the theory of electrolysis to purify and plate metals for various uses.
- The internal circuit consists of an aqueous or molten electrolyte, often with chemical additives. The external circuit consists of the electrodes, wire circuit and a power source.

7.3

- A secondary battery during recharge is an electrolytic cell, which converts electrical to chemical energy. While a battery is discharging, chemicals are reacting spontaneously, converting chemical to electrical energy, according to the theory of galvanic cells.
- For a secondary cell to recharge, it must contain a porous separator membrane, and the products of the discharge reaction must remain in contact with the electrodes so that they can be converted back into reactants.
- To reverse the chemical reactions, a higher voltage must be used than was produced by the cell when it was discharging.

7.4

- Hydrogen gas can be produced using electrolysis, PEM electrolysis or artificial photosynthesis. Current developments in this field focus on optimising energy efficiency, using minimal energy, minimising wastes and creating a 'green' hydrogen production process.

7.5

- Faraday's Laws can be used to predict the quantity of reactant and product or the amount of current or time required to consume a certain quantity of reactant or produce a certain quantity of product.
- Faraday's first law states that the charge of a cell is proportional to the mass plated at the cathode.
- Faraday's second law states that to produce 1 mol of a metal, 1, 2, 3 or another whole number of moles of electrons must be consumed.

Key formulas

Electric charge	$Q = I \times t$
Amount (in mol) of electrons	$n(\text{e}^-) = \frac{Q}{F}$
Amount (in mol)	$n = \frac{m}{M}$
Concentration	$n = c \times V$
Standard cell potential or EMF	$E^0_{\text{cell}} = E^0_{\text{reduction}} - E^0_{\text{oxidation}}$
Faraday's first law	$Q \propto m$
Faraday's second law	$Q = n(\text{e}^-) \times F$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as ‘Confidently’, ‘Mostly’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can...	Confidently	Mostly	Not really	Revision link
• describe the use and limitations of the electrochemical series to explain or predict the products of the electrolysis of chemicals, given their state (molten, liquid or aqueous) and the electrode materials used	☐	☐	☐	Go back to Topic 7.1
• write balanced equations (with states) for the reactions occurring at the anode and the cathode and the overall redox reaction for the cell	☐	☐	☐	Go back to Topic 7.1
• describe the common design features and general operating principles of commercial electrolytic cells	☐	☐	☐	Go back to Topic 7.2
• select suitable electrode materials, electrolyte (including states) and any chemical additives that result in a desired electrolysis product	☐	☐	☐	Go back to Topic 7.2
• describe the common design features and general operating principles of rechargeable (secondary) cells, with reference to discharging as a galvanic cell and recharging as an electrolytic cell, including the conditions required for the cell reactions to be reversed and the electrode polarities in each mode	☐	☐	☐	Go back to Topic 7.3
• describe the role of innovation in designing cells to meet society’s energy needs in terms of producing ‘green’ hydrogen using PEM hydrolysis and artificial photosynthesis	☐	☐	☐	Go back to Topic 7.4
• apply Faraday’s Laws and stoichiometry to determine the quantity of electrolytic reactant and product, and the current or time required to either use a particular quantity of reactant or produce a particular quantity of product	☐	☐	☐	Go back to Topic 7.5

Revision questions

Multiple choice

1 In an electrolytic cell:

A oxidation occurs at the negative electrode, which is connected to the negative terminal of the battery.

B reduction occurs at the negative electrode, which is connected to the negative terminal of the battery.

C oxidation occurs at the positive electrode, which is connected to the negative terminal of the battery.

D reduction occurs at the positive electrode, which is connected to the negative terminal of the battery.

2 In an electrolytic cell:

A the salt bridge allows for the movement of ions in the internal circuit.

B the electrolyte allows for the movement of electrons in the internal circuit.

C the salt bridge allows for the movement of electrons in the internal circuit.

D the electrolyte allows for the movement of ions in the internal circuit.

3 Compared to the voltage produced in a galvanic cell, the voltage used to recharge a cell in electrolysis is:

A lower than the voltage produced in a galvanic cell.

B equal to the voltage produced in a galvanic cell.

C greater than the voltage produced in a galvanic cell.

D always at least 1 V higher than the voltage produced in a galvanic cell.

4 A secondary cell is an electrochemical cell that can be used to produce:

A chemical energy from electrical energy by applying an external current.

B electrical energy from chemical energy when in use (discharge).

C electrical energy from chemical energy when an external current is applied.

D a constant supply of electrical energy, as there are unlimited reactants.

5 Which half-equation would be occurring at the anode of a cell while recharging?

A $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$

B $H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^-$

C $S(s) + 2H^+(aq) + 2e^- \rightarrow H_2S(g)$

D $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

6 Which half-equation would be occurring at the cathode of a cell in recharge?

A $Au(s) + 4Cl^-(aq) \rightarrow AuCl_4^-(aq) + 3e^-$

B $NO_2(g) + H_2O(l) \rightarrow NO_3^-(aq) + 2H^+(aq) + e^-$

C $MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$

D $2OH^-(aq) + H_2(g) \rightarrow 2H_2O(l) + 2e^-$

The following information refers to Questions 7 and 8.

A molten electrolytic cell was set up using inert electrodes and a solution containing a mixture of sodium chloride (NaCl) and potassium bromide (KBr).

7 Which half-equation occurs at the anode?

A $Cl_2(g) + 2e^- \rightarrow 2Cl^-(l)$

B $2Br^-(l) \rightarrow Br_2(l) + 2e^-$

C $Br_2(l) + 2e^- \rightarrow 2Br^-(l)$

D $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

8 Which half-equation occurs at the cathode?

A $Na(l) \rightarrow Na^+(l) + e^-$

B $K^+(l) + e^- \rightarrow K(l)$

C $Na^+(l) + e^- \rightarrow Na(l)$

D $K(l) \rightarrow K^+(aq) + e^-$

9 Which of the following occurs during the electrolysis of a dilute copper(II) nitrate solution using platinum electrodes?

A The negative electrode decreases in mass.

B A colourless and odourless gas is produced at the negative electrode.

C A colourless and odourless gas is produced at the positive electrode.

D The electrolyte solution becomes darker in colour.

10 During the electrolysis of a tin nitrate solution, nitrate ions are:

A attracted to the cathode, where they undergo reduction.

B attracted to the anode, where they undergo oxidation.

C not attracted to either electrode because they are spectator ions only.

D attracted to one electrode but are neither oxidised nor reduced.

Short answer

Describe and explain

11 Explain whether electrons can move through a solution.

12 Explain why the polarity of electrodes is reversed in an electrolytic cell, compared to a galvanic cell.

13 Explain why reactants are kept in the same vessel in an electrolytic cell but must be separated in a galvanic cell.

14 Explain the purpose of an electrolyte in an electrolytic cell.

15 Describe the changes experienced by a secondary cell when it finishes converting chemical to electrical energy and is then connected to a power source. Your answer must reference the following:

- reactants and products in discharge and recharge
- the requirements of a cell to recharge.

16 Explain why it is undesirable for a secondary cell to produce gases, but it is possible for a commercial electrolytic cell to do so.

17 Explain why using an iron electrode as the anode in a molten electrolytic cell containing potassium bromide (KBr) would not result in the formation of gaseous bromine ($Br_2(g)$).

18 Explain why using a copper electrode as the anode in an aqueous electrolytic cell containing a solution of sodium chloride (NaCl) would not result in the cell forming a gaseous product.

19 Explain why a calcium fluoride (CaF_2) electrolytic cell must be molten and cannot be conducted in aqueous conditions.

20 Explain why it is preferred for a chlorine (Cl_2) gas-producing iron(II) chloride ($FeCl_2$) electrolytic cell to use aqueous reactants at non-standard conditions, rather than molten reactants.

Apply, analyse and compare

21 Magnesium has many purposes in industry. Nearly 70% of all magnesium produced is used in metal alloys due to its strength and low density. Magnesium is produced by electrolysis of molten magnesium chloride ($MgCl_2$) with some impurities using the following electrochemical cell.

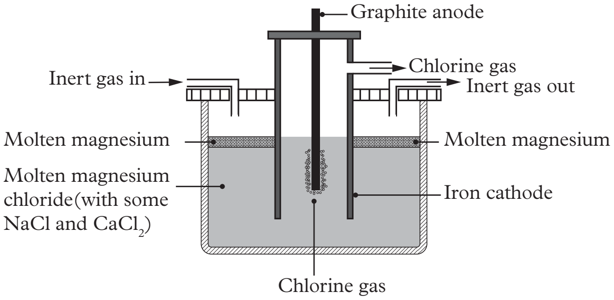


FIGURE 1 A magnesium electrolytic cell

Source: *VCE Chemistry 2014 Chemistry Exam, Question 9* reproduced by permission © VCAA

a Construct the two half-equations for the reactions occurring at the anode and the cathode.

b Construct a balanced overall reaction equation.

c Explain, with the use of half-equations, why the sodium and calcium impurities are acceptable in the molten electrolyte.

d Explain, with the use of half-equations, what would happen if iron(II) chloride ($FeCl_2$) was one of the impurities.

e Using half-equations and a balanced overall equation, explain what would happen if the cell was aqueous, rather than molten.

22 Blister copper is purified using the following electrochemical cell, which uses an aqueous acidified copper(II) sulfate ($CuSO_4$) electrolyte.

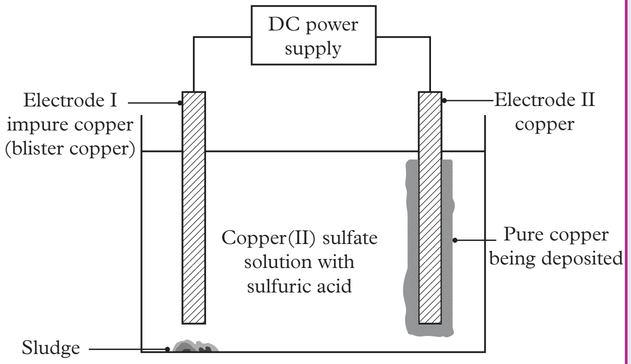


FIGURE 2 An electrolytic cell for the purification of copper from blister copper

Source: *VCE Chemistry 2015 Chemistry Exam, Question 28–30* reproduced by permission © VCAA

a Identify the anode and the cathode (when the power supply is turned on).

b Construct the two half-equations for the reactions occurring at the anode and the cathode.

c Explain, with the use of half-equations, why silver and gold are collected as by-products of the process.

d Using half-equations and a balanced overall equation, explain what would happen if the cell was run with a molten copper(II) sulfate ($CuSO_4$) electrolyte, rather than an aqueous one.

23 A secondary battery uses the following half-equations:

$Fe(OH)_2(s) + 2e^- \rightarrow Fe(s) + 2OH^-(aq)$ E^0 = positive

$Ni(OH)_3(s) + e^- \rightarrow Ni(OH)_2(s) + OH^-(aq)$ E^0 = negative

a Identify the anode and cathode half-equations in both discharge and recharge, and determine the resulting polarity of electrodes.

b Construct the overall equations for the reactions occurring in discharge and recharge modes.

c Describe the nature of the electrolyte in this cell.

24 Nickel metal is plated onto an iron cup using an aqueous electrolyte of nickel(II) nitrate ($Ni(NO_3)_2$). Draw the electrolytic cell for this process, labelling:

- the anode and the cathode
- electrode polarity
- the electrode materials
- oxidation and reduction half-equations
- the strongest oxidising and reducing agents
- the overall equation
- the direction of ion movement
- the direction of electron movement
- the voltage required to operate the cell.

25 For each of the following, construct an electrolytic cell that would result in these outcomes. You must include the information listed in Question 24.

a $H_2(g)$ forms at a negative cathode.

b $Cl_2(g)$ forms at a positive anode.

c Ag is plated onto a negative cathode.

d The pH of the electrolyte in the cell decreases.

26 Construct the following electrolytic cells, which contain:

a a mixture of molten lithium chloride (LiCl) and molten calcium iodide (CaI_2)

b a mixture of aqueous calcium nitrate ($Ca(NO_3)_2$), aqueous cadmium bromide ($CdBr_2$) and aqueous potassium chloride (KCl)

c an aqueous mixture of tin(IV) chloride ($SnCl_4$) and lithium iodide (LiI).

You must include the information listed in Question 24.

27 An aqueous sodium bromide (NaBr) electroplating cell was made using 1.0 M NaBr. Both electrodes are made of graphite. A voltage of 12.0 V is applied, creating a current of 4.90 A for 13.50 minutes. Calculate the mass of product generated at the cathode in this time.

28 A molten electrolysis cell consisting of a mixture of manganese chloride ($MnCl_2$) and nickel(II) iodide (NiI_2) was run with a current of 100 A for 3.0 hours. The electrodes are made of platinum and products are continually removed. Calculate the mass of product generated at the cathode every hour.

29 In an electrolytic cell, 1.500 L of a 1.00 M hydrofluoric acid (HF) solution undergoes electrolysis with a current of 12.0 A for 30.0 minutes. Both electrodes are made of graphite.

a Calculate the volume of product generated at the cathode at SLC.

b Calculate the new concentration of the HF solution after 30 minutes.

30 To plate 5.20 g of cobalt onto the cathode of a cell, a current of 13.0 A is applied to an aqueous solution of cobalt(II) iodide (CoI_2). Calculate how long (in minutes) the cell has to operate to deposit this mass onto the electrode.

31 A scientist has 1.00 hour to plate 100 g of gold onto the cathode of a cell using a molten salt of gold fluoride (AuF). Calculate the current required to ensure the mass is plated in this time frame.

32 An industrial scientist has been challenged to produce 1.00 kg of aluminium per hour in a 24-hour period using the Hall–Héroult cell. Calculate the current required per hour to ensure the mass is plated in the 24 hours.

33 A current of 8.5 A is run for 15 minutes to plate a mass of 5.03 g of copper at the cathode of a cell. Determine the charge on the copper ion.

34 A current of 5.48 A is run for 30 minutes, plating a mass of 2.85 g of iron at the cathode of a cell. Determine the charge on the iron ion.

Design and discuss

35 The temperature of the human body is slightly higher than 37°C. Investigate the impact of running galvanic electrolysis, to kill hair follicles, at this temperature. Discuss the chemical half-equations that would result under these non-standard conditions and make a conclusion about the viability of this method at body temperature.

36 Discuss the products, using half-equations and an overall equation, that would result from electrolytic cells where the electrolyte is:

a molten NaCl

b concentrated NaCl

c dilute NaCl.

37 Use a mind map to communicate the links and ideas between the following concepts:

redox, oxidation, reduction, oxidising agent, reducing agent, galvanic cell, primary cell, electrolytic cell, secondary cell, fuel cell, salt bridge, electrolyte, anode, cathode, polarity, E^0 and power source

38 Summarise chapters 4 and 7 by discussing the similarities and differences between galvanic cells, primary cells, secondary cells and fuel cells. Your summary must include:

- structural and chemical differences
- structural and chemical similarities
- advantages
- disadvantages
- applications.

39 The mercuric oxide battery was widely used until its ban in 1991. It was capable of producing a constant voltage of 1.35 V and had many applications. Investigate the mercuric oxide battery, then answer the following questions.

a Construct half-equations for the reactions occurring at the cathode and anode in discharge, labelling them as oxidation or reduction.

b Construct the overall reaction equation for the reaction occurring in discharge.

c Use these equations to evaluate the battery, specifically communicating why it was banned in 1991.

d Investigate the health effects this type of battery had on users.

e Describe the battery that was introduced as a substitute for the mercuric oxide battery.

f Construct the half-equations for the reactions occurring at the cathode and anode in discharge of the substitute battery, labelling them as oxidation or reduction.

g Construct the overall reaction equation for the reaction occurring in the substitute battery in discharge.

h Use these equations to evaluate the battery, specifically communicating why it is safer than the mercuric oxide battery.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

AREA OF STUDY 2

Checkpoint

Multiple choice

Question 1

In endothermic reactions:

A the activation energy is lower than the enthalpy.

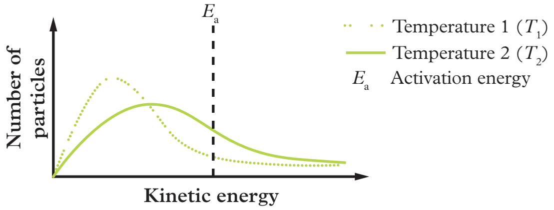
B the energy of the reactants is higher than the energy of the products.

C the addition of a catalyst will decrease the enthalpy.

D the energy absorbed to break the reactants' bonds is greater than the energy released when products are formed.

Question 2

Consider the following distribution of the kinetic energy of particles in a reaction system at two different temperatures.



A $T_1 > T_2$. At T_1, more particles have energy equal to or greater than the activation energy and, therefore, there is a greater proportion of successful collisions.

B $T_1 > T_2$. At T_2, more particles have energy equal to or greater than the activation energy and, therefore, there is a greater proportion of successful collisions.

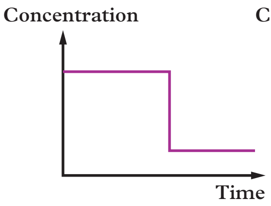
C $T_2 > T_1$. At T_1, more particles have energy equal to or greater than the activation energy and, therefore, there is a greater proportion of successful collisions.

D $T_2 > T_1$. At T_2, more particles have energy equal to or greater than the activation energy and, therefore, there is a greater proportion of successful collisions.

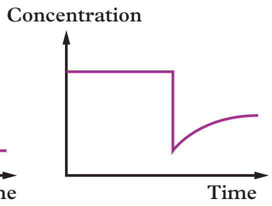
Question 3

Consider the following equilibrium system:

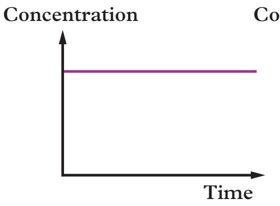
$$Cu^{2+}(aq) + 4Cl^{-}(aq) \rightleftharpoons CuCl_4^{2-}(aq)$$



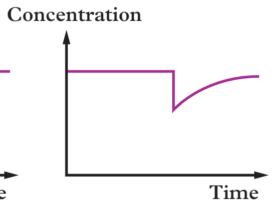
Graph I



Graph II



Graph III



Graph IV

When the volume of the system is halved, which graph represents the change experienced by the Cu^{2+} ions?

A Graph I

B Graph II

C Graph III

D Graph IV

Question 4

Nitrogen and hydrogen react to form ammonia according to the following equilibrium equation:

$$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$$

15.00 mol of ammonia are pumped into an empty 5.00 L vessel and allowed to reach equilibrium. At equilibrium, the concentration of nitrogen is found to be 1.2 M. Calculate the equilibrium constant of the reaction.

A $6.43 \times 10^{-3}\,M^{-2}$

B $155.52\,M^{-2}$

C $7.52 \times 10^{-3}\,M^{-2}$

D $2.25\,M^{-2}$

Question 5

Which of the following will affect the position of equilibrium?

A The addition of a catalyst

B The addition of argon gas

C The removal of a product

D None of the above

Question 6

Which of the following accurately depicts the polarity of the electrodes in galvanic and electrolytic cells?

	Galvanic cell		Electrolytic cell	
	Anode	Cathode	Anode	Cathode
A	+	–	–	+
B	+	–	+	–
C	–	+	+	–
D	–	+	–	+

Question 7

In an electrolytic cell, the purpose of the electrolyte is to:

A allow ion movement between the anode and the cathode.

B allow electron movement between the anode and the cathode.

C contain water to ensure that all salts remain in solution in all electrolytic cells.

D prevent electrons from moving from the anode to the cathode.

Question 8

Which of the following statements is accurate about molten electrolytic cells?

A Molten cells are always conducted at standard laboratory conditions.

B Positive ions in the electrolyte are attracted to the negative electrode.

C Molten mixtures will increase the temperature required to liquify the electrolyte.

D Water in the electrolyte is attracted to the positive electrode.

Question 9

In an electrolytic cell, 150.0 mL of a 1.2 M hydrochloric acid (HCl) solution undergoes electrolysis with a current of 12.0 A for 10.0 min. Determine the new concentration of the HCl solution.

A 0.50 M

B 0.70 M

C 8.3×10^{-3} M

D 1.19 M

Question 10

A current of 21.3 A is applied for 5.50 min to plate a mass of 0.74 g of vanadium at the cathode of a cell. Determine the charge on the vanadium ion.

A +2

B +3

C +4

D +5

Short answer

Question 1 (6 marks)

Hydrogen and bromine react to form hydrogen bromide. The overall reaction proceeds according to the following equation:

$$H_2(g) + Br_2(g) \rightleftharpoons 2HBr(g) \qquad \Delta H = \text{negative}$$

Bromine gas has a reddish-brown colour, whereas hydrogen and hydrogen bromide are colourless. Considering the system at equilibrium, predict the effect of each of the following changes on the percentage yield of hydrogen bromide and give your reasoning.

a The volume of the system is doubled. 2 marks

b The temperature of the system is increased. 2 marks

Predict the effect of the following change on the final colour of the gas mixture, if its initial colour is light brown.

c Bromine is added into the system. 2 marks

Question 2 (7 marks)

Nitrosyl chloride (NOCl) is a red gas that acts as a catalyst and a bleaching agent. It will decompose into NO, a colourless gas, and Cl_2, a toxic yellow gas.

$$2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g) \qquad \Delta H = \text{positive}$$

$$K = 1.6 \times 10^{-5}\ \text{M at } 25°\text{C}$$

a Initially, an empty 3.0 L reaction vessel is filled with 6.0 mol of NO and 12 mol of Cl_2. Approximately 30 seconds into the reaction, the concentration of NOCl is found to be 1.2 M. Calculate the reaction quotient, Q, at this time. 3 marks

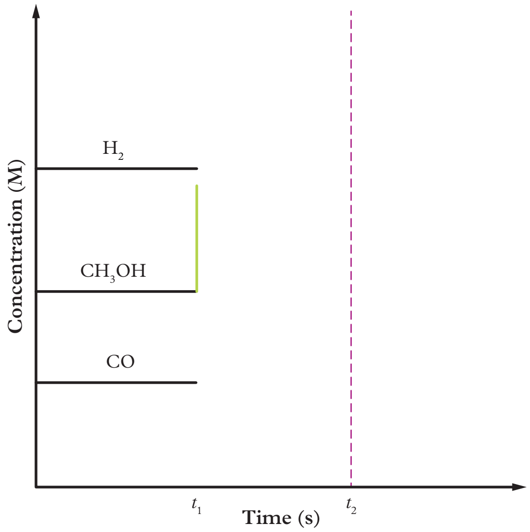
b Compare the K and Q values to determine the direction in which the system must shift to reach equilibrium. Explain your response. 2 marks

c The volume of the container was doubled. Describe the final colour of the system once equilibrium is re-established. Explain your answer. 2 marks

Question 3 (9 marks)

Methanol, a liquid used as a car fuel additive, is commercially produced from carbon monoxide and hydrogen at 300°C according to the following reaction:

$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g) \quad \Delta H = -92\ J\,mol^{-1}$



a On the graph, draw the concentration profile that results from an increase in concentration of methanol at t_1, if equilibrium is re-established at t_2. 2 marks

b At equilibrium, the amounts of carbon monoxide, hydrogen and methanol in a 4 L vessel are found to be 1.2 mol, 0.5 mol and 4.6 mol respectively.

- **i** Write the equilibrium expression for the reaction. 1 mark
- **ii** Calculate the equilibrium constant. 2 marks

c In an attempt to increase the rate and yield of the reaction, a chemist increases the temperature to 400°C. Explain the effect of this change on the rate and yield of methanol production. 4 marks

Question 4 (5 marks)

Artificial photosynthesis uses a water oxidation and proton reduction catalyst system to split water into oxygen and hydrogen gases.

a Construct the half-equation for the water oxidation reaction. 1 mark

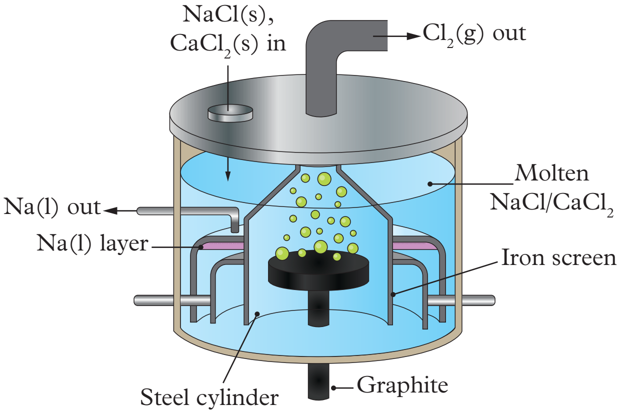
b Construct the half-equation for the proton reduction reaction. 1 mark

c Draw an arrow (→ or ←) to identify the direction of flow of electrons. 1 mark

d Explain how artificial photosynthesis is being used as an alternative method to produce 'green' fuel. 2 marks

Question 5 (7 marks)

The Downs cell is used to electrolyse a mixture of molten NaCl and $CaCl_2$ at 600°C to produce pure sodium metal. Sodium metal has a lower density than the mixture and therefore rises to the top of the container. This ensures it is separated from the gas produced at the graphite electrode. It then cools and is immediately stored within oil.



a Construct the half-equation occurring at the graphite electrode. 1 mark

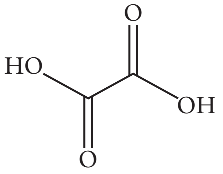
b State the polarity of the steel container. 1 mark

c Explain, using chemical equations, why the sodium must be separated from the gaseous product. 2 marks

d Calculate the mass, in tonnes, of sodium metal produced when the cell is run at 6.00 V and 30 000 A over a period of 48.0 hours. 3 marks

Question 6 (6 marks)

A student develops an experiment to determine the effect of the surface area on the rate of a reaction. Rhubarb sticks containing oxalic acid (shown below) are cut into cubes with surface areas of 0.5 cm^3, 1 cm^3 and 2 cm^3. The cubes are used to reduce an acidified permanganate solution and form carbon dioxide gas.



The rhubarb cubes are placed in a beaker of the pale pink permanganate solution and the results are recorded as the time taken for the solution to turn clear. The results are outlined in the table.

Surface area (cm^3)	Time (minutes)
0.5	0.8
1.0	2.1
2.0	3.2

a Identify the independent variable. 1 mark

b Identify the dependent variable. 1 mark

c Write a conclusion based on the data given in the table. 1 mark

d Comment on the precision of the experiment. 1 mark

e Identify a variable that has not been controlled. 1 mark

f Explain how the variable in part **e** affects the validity of the experiment. 1 mark

TOTAL MARKS

/50 marks

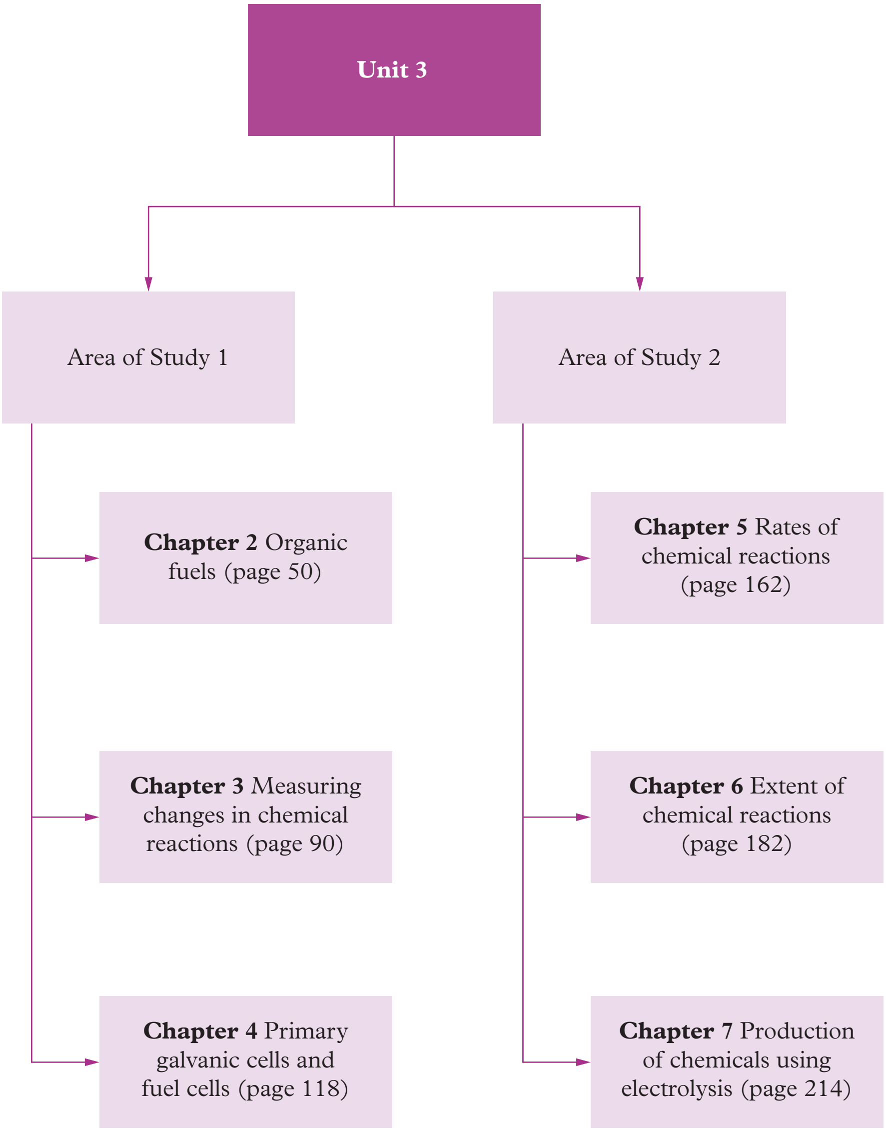
UNIT 3 Review

PART A – Revisit and revise

Part A of the Unit Review will help you revisit and revise all the key concepts and terms from Unit 3 and test your understanding to identify strengths and weaknesses.

Unit 3 overview

The chart below shows all the Areas of Study for Unit 3 and the relevant chapters in your student book. Go to the pages shown to review the key concepts for each chapter.



Test your understanding

Use the following table to guide your revision:

Step 1 – Read the key knowledge for this unit.

Step 2 – Test your understanding of the key knowledge by answering the question(s).

Step 3 – Rate your understanding of each key knowledge from low to high.

Step 4 – Use the topic and page numbers to revise the concepts for which you've identified you need practice.

Only the first four key knowledge dot points are shown. Access the rest of the Test your understanding questions in your **obook pro**.

Key knowledge	Test yourself	Rate yourself	Focus your revision
• the definition of a fuel, including the distinction between fossil fuels (coal, natural gas, petrol) and biofuels (biogas, bioethanol, biodiesel) with reference to their renewability (ability of a resource to be replaced by natural processes within a relatively short period of time)	1 Describe and explain the differences in composition between diesel and biodiesel fuels.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 2.1 Pages 54–57
	2 Compare and contrast fossil fuels and biofuels in terms of how they are formed and their renewability.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 2.1 Pages 54–57
• fuel sources for the body measured in $kJ\,g^{-1}$: carbohydrates, proteins and lipids (fats and oils)	3 The energy content of a muesli bar is 1860 kJ per 100 g. If the recommended serving size of 1 bar is 45 g, how much energy is consumed per serve?	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 2.2 Pages 58–61
• photosynthesis as the process that converts light energy into chemical energy and as a source of glucose and oxygen for respiration in living things: $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$	4 Using a balanced chemical equation, explain why photosynthesis is important and where it occurs.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 2.3 Pages 62–63
• oxidation of glucose as the primary carbohydrate energy source, including the balanced equation for cellular respiration: $C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$	5 Write the balanced chemical equation for the oxidation of glucose and explain the importance of this process.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 2.4 Pages 64–66

PART B – Exam essentials

Now that you've completed your revision for Unit 3, it's time to learn and practise some of the skills you'll need to answer exam questions like a pro! To help you, our expert authors have created the following advice and tips to help you maximise your results on the end-of-year examination.

Exam tip 1: Keep moving forward

- If your values look odd, there is a chance that you have made a mistake.
- Do not go back and look for where you went wrong … unless you know exactly where you made the mistake, or you have time at the end of the exam. So many students do not complete their exam because they have wasted time looking for where they made mistakes.
- The examiners will only deduct marks where you made the mistake. They will recalculate all your remaining answers so that if you have used the incorrect value in the correct way, you will still be awarded full marks in the next parts of the question. These are called consequential marks.

See it in action

Read the real exam question below and see how the tip has made a difference in the high-scoring and low-scoring responses.

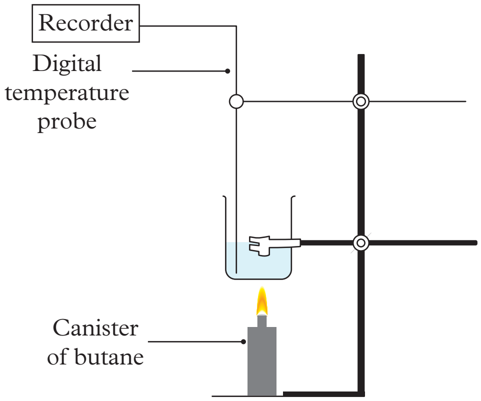
Question 10

A senior Chemistry student bought a packet of Krispy Krackers biscuits from the local farmers' market. The label on the packet had no information on the energy content of the biscuits.

The student decided that he would measure the energy content of a Krispy Krackers biscuit by burning it under a can of water and measuring the temperature rise of the water.

Having performed a similar experiment in class, he knew that when the biscuit was burnt, heat energy would be lost to the environment. To increase the accuracy of the result, some butane gas from a butane canister was first burnt and the temperature rise of the water from that was measured. The heat energy released by burning butane was known, and the percentage energy loss using the equipment could be determined and adjusted to get the result for the biscuit.

The experimental set-up and the results for Part 1 of the experiment are shown below.



Part 1 – The heat content of butane

1. Measure the initial mass of a butane canister.
2. Measure the mass of a metal can, add 250 mL of water and re-weigh.
3. Set up the apparatus as in the diagram and measure the initial temperature of the water.
4. Burn the butane gas for 5 minutes.
5. Immediately measure the final temperature of the water.
6. Measure the final mass of the butane canister when cool.

Results table for Part 1

Quantity	Measurement
mass of empty can	52.14 g
mass of can + water before combustion	303.37 g
mass of butane canister before heating	260.15 g
mass of butane canister after heating	259.79 g
initial temperature of water	22.1°C
final temperature of water	32.7°C

a. ii. Calculate the amount of heat energy absorbed by the water when it was heated by burning the butane. Give your answer in kilojoules. 2 marks

iii. Calculate the experimental value of the molar heat of combustion of butane. Give your answer in kJ mol^{-1}. 2 marks

Source: *VCE 2016 Chemistry Exam* reproduced by permission © VCAA

A high-scoring response

a. ii. Calculate the amount of heat energy absorbed by the water when it was heated by burning the butane. Give your answer in kilojoules.

$m(H_2O) = 303.37 - 52.14$

$= 251.23\ g$

$\Delta T(H_2O) = 32.7 - 22.1$

$= 10.6°C$

$q = mc\Delta T$

$= 251.23 \times 4.18 \times 10.6$

$= 11131\ J \div 1000 = 11.1\ kJ$

2 marks

✓

Has identified that the answer should be given in kilojoules and has done the conversion

iii. Calculate the experimental value of the molar heat of combustion of butane. Give your answer in kJ mol^{-1}.

$m(C_4H_{10})$ consumed/reacted $= 260.15 - 259.79$

$= 0.36\ g$

$n(C_4H_{10}) = \frac{m}{M} = \frac{0.36}{58.0}$

$= 0.0062\ mol$

$\Delta H = \frac{q}{n} = \frac{11.1}{0.0062}$

$= 1793.41\ kJ\ mol^{-1} = -1.8 \times 10^3\ kJ\ mol^{-1}$

2 marks

Correct answer provided, including the – sign for the Δ*H*, as it is an exothermic combustion reaction.

Heat energy in kilojoules from **ii**

A low-scoring response

a. ii. Calculate the amount of heat energy absorbed by the water when it was heated by burning the butane. Give your answer in kilojoules.

$m(H_2O) = 303.37 - 52.14$

$= 251.23$ g

$\Delta T(H_2O) = 32.7 - 22.1$

$= 10.6$°C

$q = mc\Delta T$

$= 251.23 \times 4.18 \times 10.6$

$= 11131$ J

2 marks

✗

Calculation was correct, but the student did not convert joules to kilojoules

iii. Calculate the experimental value of the molar heat of combustion of butane. Give your answer in kJ mol^{-1}.

$m(C_4H_{10})$ consumed/reacted $= 260.15 - 259.79$

$= 0.36$ g

$n(C_4H_{10}) = \frac{m}{M} = \frac{0.36}{58}$

$= 0.0062$ mol

$\Delta H = \frac{q}{n} = \frac{11131}{0.0062}$

$= 1793408$ kJ mol^{-1} $= -1.8 \times 10^6$ kJ mol^{-1}

2 marks

Answer is incorrect since the student used the heat energy in joules, instead of heat energy in kilojoules. They performed the correct calculation with the wrong value, so they will still be awarded one mark.

Think like an examiner

To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam. Imagine you are an examiner and use the marking guidance below to mark the response.

Question 2

A vehicle that is powered by a diesel engine is able to use either petrodiesel or biodiesel as a fuel. Petrodiesel and biodiesel are not pure substances, but are a mixture of molecules. In general, petrodiesel consists of molecules that are shorter in length, on average, than those found in biodiesel. Biodiesel contains molecules that include functional groups.

The table below lists some of the properties of the two fuels.

Fuel	Major component	Energy content ($MJ\,kg^{-1}$)	CO_2 emission ($kg\ CO_2/kg$ of fuel)
petrodiesel	$C_{12}H_{26}$	43	3.17
biodiesel	$C_{19}H_{32}O_2$	38	2.52

b. Assume that combustion occurs in an unlimited supply of oxygen for each of the following calculations.

i. Calculate the number of litres of biodiesel that are required to be burnt to produce the same amount of energy as 2.5 kg of petrodiesel (density (biodiesel) = $0.89\,kg\,L^{-1}$)

Energy in petrodiesel = 2.5 × 43

= 107.5 MJ

m(biodiesel) = 107.5 ÷ 38 = 2.8 kg

V(biodiesel) = 0.89 × 2.8 = 2.52 mL

3 mark

ii. Calculate the mass of carbon dioxide, CO_2, that would be produced from 3.91 kg of biodiesel.

$m(CO_2) = 3.91 \times 3.17 = 12.39$ kg

1 mark

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Marking guide

Question 2bi	1 mark for correctly calculating the energy in 2.5 kg of petrodiesel
	1 mark for correctly calculating the mass of the biodiesel
	1 mark for correctly calculating the volume of the biodiesel
Question 2bii	1 mark for calculating the mass of the CO_2

Note: In chemistry, you can be assigned consequential marks. This is where, if an answer to a previous question is required for another response (i.e. **1b** is needed for **1c**), and you get the first part wrong, assessors will complete the calculation using the incorrect value. This ensures that students are not penalised for incorrect values twice.

Fix the response

Consider where you did and did not award marks in the above response. How could the response be improved?

Write your own response to the same question to receive full marks from an examiner.

Exam tip 2: Identify and use key terms in your responses

- Exam questions may ask you to use your understanding of chemistry to explain how phenomena occur or predict the effect of changes. Make sure to identify the relevant key terms and use them in your answer.
- For example, when explaining the effect on the position of equilibrium, a key term to include is 'partially oppose'. Your answer should also indicate the 'net' reaction being forward or backward/reverse.

See it in action

Read the real exam question below and see how the tip has made a difference in the high-scoring and low-scoring responses.

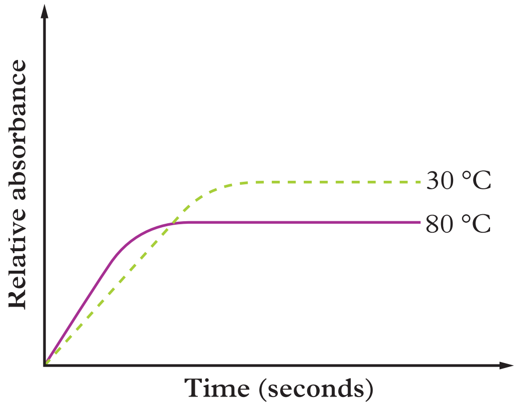
Question 3

The cobalt(II) tetrachloride ion, $CoCl_4^{2-}$, dissociated into the cobalt(II) ion, Co^{2+}, and chloride ions, Cl^-, according to the following chemical equation.

$$\underset{\text{blue}}{CoCl_4^{2-}(aq)} \rightleftharpoons \underset{\text{pink}}{Co^{2+}(aq)} + 4\,Cl^-(aq)$$

20 mL samples of the equilibrium mixture were heated to two temperatures, 30°C and 80°C. The intensity of the pink colour of the Co^{2+} product was recorded every 30 seconds by measuring the absorbance of the solution. The higher the intensity of the pink colour, the higher the absorbance.

The results of this experiment are shown in the graph below.



a. State whether the forward reaction is exothermic or endothermic. Justify your answer by referring to the graph.

2 marks

Source: *VCE 2019 Chemistry Exam* reproduced by permission © VCAA

A high-scoring response

a. State whether the forward reaction is exothermic or endothermic. Justify your answer by referring to the graph.

Has used the graph to draw this conclusion

Provided the correct answer

✓ The relative absorbance of pink colour is higher at lower temperatures, meaning there is a higher concentration of the product Co^{2+}/Cl^-. The reaction is exothermic, as it will partially oppose a decrease in temperature by favouring a net forward reaction to increase the concentration of Co^{2+}/Cl^-, resulting in a higher intensity of pink.

2 marks

Used the relevant key terms in their justification

A low-scoring response

a. State whether the forward reaction is exothermic or endothermic. Justify your answer by referring to the graph.

Did use a key term, but has provided no justification

✗ The reaction is endothermic because there is a net forward reaction.

2 marks

Provided an incorrect answer. They have not referred to the graph, which would have helped them think through this question step by step.

Think like an examiner

To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam. Imagine you are an examiner and use the marking guidance below to mark the response.

Question 8

The reaction of the oxidation of sulfur dioxide, SO_2, is shown below:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H = -197\,kJ\,mol^{-1}$$

c. The volume of the closed container is doubled.

Describe the effect that this has on the concentration of SO_2 from the time just before the volume was changed until after the system has re-established its equilibrium.

As the volume increases, the concentration of all reactants and products decreases. The system opposes the change by moving to the side with more particles, resulting in a forward reaction. This increases the concentration of SO_2. Therefore, the concentration of SO_2 is higher than it originally was before the change.

3 marks

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Marking guide

Question 8c	1 mark for stating that an increase in volume will decrease the concentration of SO_2 (not everything, as you are only being asked about SO_2)
	1 mark for stating that the system ‘partially opposes’ the change by ‘favouring the side with more particles’ and resulting in a ‘net forward’ reaction, which increases the concentration of SO_2
	1 mark for stating that, once equilibrium is re-established, the concentration of SO_2 is lower than the original concentration (before the volume was increased)

Fix the response

Consider where you did and did not award marks in the above response. How could the response be improved?

Write your own response to the same question to receive full marks from an examiner.

Practice makes perfect

Now that you know all these tips, it’s time for you to move on to Part C – Exam practice, to put them into practice.

PART C – Exam practice

Multiple choice

Question 1

A sample of NOCl is allowed to reach equilibrium according to the following equation:

$$2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$$

The volume of the mixture is halved and allowed to reach a new equilibrium, with the temperature remaining constant. What happens to the chlorine at the new equilibrium?

A It's decreased in amount and decreased in concentration.

B It's increased in amount and decreased in concentration.

C It's decreased in amount and increased in concentration.

D It's increased in amount and increased in concentration.

Question 2

Carbon dioxide gas dissolves to a small extent in water, forming carbonic acid in an exothermic reaction:

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \qquad \Delta H = \text{negative}$$

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$$

This is the reaction involved in forming carbonated or fizzy drinks.

Which of the following would *not* be effective at increasing the amount of dissolved carbon dioxide?

A Decreasing the pH of the solution

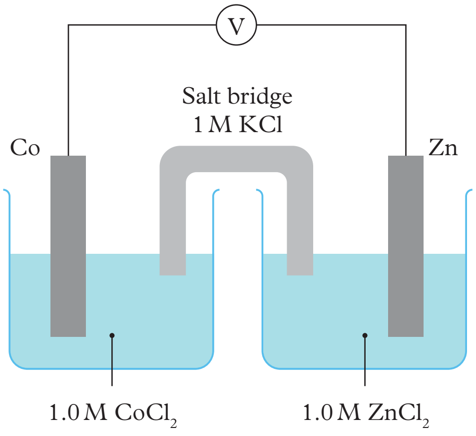
B Decreasing the temperature of the solution

C Increasing the concentration of carbon dioxide in the gas

D Increasing the pressure of the carbon dioxide gas

Question 3

An electrochemical cell is set up as shown.



Calculate the initial voltage of the cell.

A 0.48 V

B −0.48 V

C 0.00 V

D 1.04 V

Question 4

Samples of uranium, vanadium and yttrium (U, V, Y) were placed in solutions of the metallic ions U^{3+}, V^{2+} and Y^{3+}. The following observations were recorded:

Trial	Ion	Metal	Observation
1	U^{3+}	Y	reaction
2	V^{2+}	U	reaction
3	V^{2+}	Y	reaction
4	Y^{3+}	V	no reaction

The order of oxidising agents from the strongest to the weakest is:

A V^{2+}, U^{3+}, Y^{3+}

B U^{3+}, V^{2+}, Y^{3+}

C Y^{3+}, U^{3+}, V^{2+}

D V^{2+}, Y^{3+}, U^{3+}

Question 5

Camping stoves use propane gas, which burns according to the equation:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l) \qquad \Delta H = -2217\ \text{kJ mol}^{-1}$$

A particular stove had a heat output of 22.17 kJ per minute. For each minute of operation:

A 0.05 g of O_2 is used.

B 0.44 g of propane is used.

C 1.8×10^{24} molecules of CO_2 are produced.

D 0.72 kg of H_2O is produced.

Question 6

In the electroplating of silver, 0.2 faraday will:

A be equal to 1 amp of current flowing for 5 seconds.

B be equal to 0.2 mol of electrons per second.

C reduce 0.1 mol of Ag^+ ions.

D produce about 21.6 g of silver.

Question 7

The overall reaction during the discharge of a lead–acid accumulator is:

$$Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{\ 2-}(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$$

When the battery is recharging:

A the pH of the electrolyte increases.

B all Pb^{2+} ions are reduced to Pb atoms.

C Pb^{2+} ions are either oxidised or reduced, depending on which electrode they are near.

D the lead plates of the battery decrease in mass.

Question 8

An experiment was used to determine the energy available from a biscuit. The biscuit was burnt and the energy released was used to heat 250 g of water. When a 0.657 g piece of biscuit was burnt, the temperature of the 250 g of water increased by 10.9°C. What was the energy released by this biscuit, assuming all of the energy was absorbed by the water?

A 7.48 kJ

B 11.4 kJ

C 41.8 kJ

D 69.3 kJ

Question 9

The energy released when four chemicals, which can be used as fuels, undergo complete combustion is given in the table below:

Fuel	Energy released (kJ mol^{-1})
Methane (CH_4)	803
Ethanol (C_2H_5OH)	1235
Propane (C_3H_8)	2045
Octane (C_8H_{18})	5077

Which of these fuels releases the most energy for each mole of carbon dioxide produced?

A Methane

B Ethanol

C Propane

D Octane

Question 10

Two alternative power sources being developed for motor vehicles are new secondary cells and fuel cells. Which of the following statements best describes the difference between a fuel cell and a secondary cell?

A Fuel cells can be recharged by the application of an external power supply while secondary cells require fresh reactants to be continuously supplied to the cell producing the electricity.

B Fuel cells do not produce any waste materials whereas secondary cells produce some waste materials that can be hazardous.

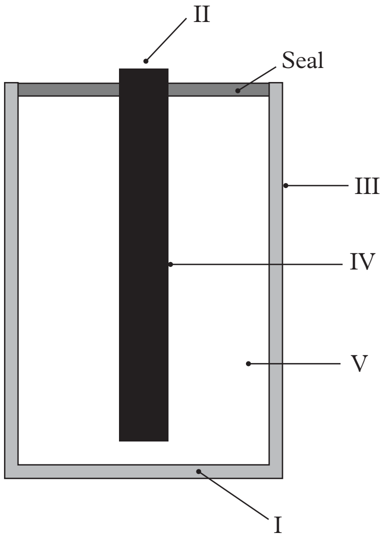
C Secondary cells can be easily recharged whereas fuel cells cannot be recharged.

D Secondary cells can be recharged by the application of an external power supply while fuel cells require fresh reactants to be continuously supplied to the cell producing the electricity.

Short answer

Question 1 (7 marks)

A cross-sectional diagram of the components of a standard non-rechargeable galvanic cell is given:



a Identify the polarity of electrode I. (1 mark)

b Identify the polarity of electrode IV. (1 mark)

c Suggest materials that could be used for the outside container of the cell and describe its role in the cell. (2 marks)

d Suggest materials that could be used for electrode IV. (1 mark)

e Suggest materials for the composition and role of material V. (2 marks)

Question 2 (5 marks)

The combustion of graphite has the following thermochemical equation.

$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -393\,kJ\,mol^{-1}$

When 0.12 g of graphite is burnt in a calorimeter, the temperature rises from 20.82°C to 22.06°C.

a Use this information to calculate the calibration factor for the calorimeter. (3 marks)

b 1.8 g of glucose was then burnt in the calorimeter, according to the equation:

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

The temperature rose further to 31.06°C.

Calculate ΔH for the reaction. (2 marks)

Question 3 (10 marks)

The equilibrium constant for a chemical reaction occurring at a fixed temperature with gaseous reactants and products can be expressed as:

$$K = \frac{[A]^2[B]}{[C][D]^3} = 7.46 \times 10^5\,M^{-1}$$

a Construct the balanced chemical equation for the reaction involving gases A, B, C and D. (1 mark)

b Use the value of the equilibrium constant to determine whether the equilibrium mixture will contain only reactants, only products, or significant amounts of both reactants and products. (1 mark)

c Deduce the ratio of reactants to products when this reaction mixture is at equilibrium. (1 mark)

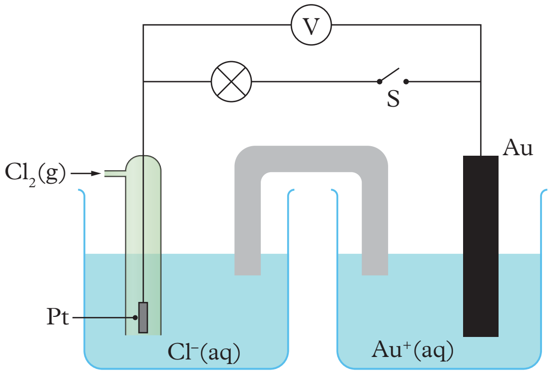
d If some additional gas C is added to the equilibrium mixture, state the effect on the value of *K*. (1 mark)

e If this reaction is exothermic, describe and explain the effect on the value of *K* when the temperature is increased. (2 marks)

f Under a different set of temperature conditions, a 2.0 L container with 0.5 mol of gas C and 0.75 mol of gas D is allowed to reach equilibrium. The amount of gas B at equilibrium is 0.15 mol. Calculate the value of *K* at this temperature. (4 marks)

Question 4 (9 marks)

The diagram below shows an electrochemical cell operating under standard conditions.



a Construct the equations occurring at the anode and cathode and the overall equation for the cell. (3 marks)

b When the switch S is open, calculate the voltage of the cell. (1 mark)

c Identify the oxidising agent. (1 mark)

d Calculate the standard reduction potential for Au, if the E^0_{cell} is +0.32 V. (2 marks)

e Identify suitable material(s) that could be used to construct the salt bridge for this cell and give reasons for your choice. (2 marks)

Question 5 (4 marks)

The following reaction takes place in a lead–acid storage battery:

$$2PbSO_4(s) + 2H_2O(l) \underset{\text{Discharge}}{\overset{\text{Charge}}{\rightleftharpoons}} Pb(s) + PbO_2(s) + 2H_2SO_4(aq)$$

a Construct the balanced half-equations for the reactions at the anode and the cathode during the charging process. (2 marks)

b Propose one advantage and one disadvantage of a lead–acid storage battery compared to a zinc–carbon battery. (2 marks)

Question 6 (2 marks)

The thermochemical equations for the combustion of glucose and a fatty acid are given below.

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) \quad \Delta H = -2803\ \text{kJ mol}^{-1}$$

$$C_{16}H_{32}O_2(aq) + 23O_2(g) \rightarrow 16CO_2(g) + 16H_2O(l) \quad \Delta H = -9770\ \text{kJ mol}^{-1}$$

Using the above information, prove that a weight of fats is more fattening than the same weight of carbohydrates.

Question 7 (3 marks)

Reacting methane and water is one way to prepare hydrogen:

$$CH_4(g) + H_2O(l) \rightarrow CO(g) + 3H_2(g)$$

If you begin with 995.0 g of CH_4 and 2510 g of water, determine the maximum possible yield of H_2.

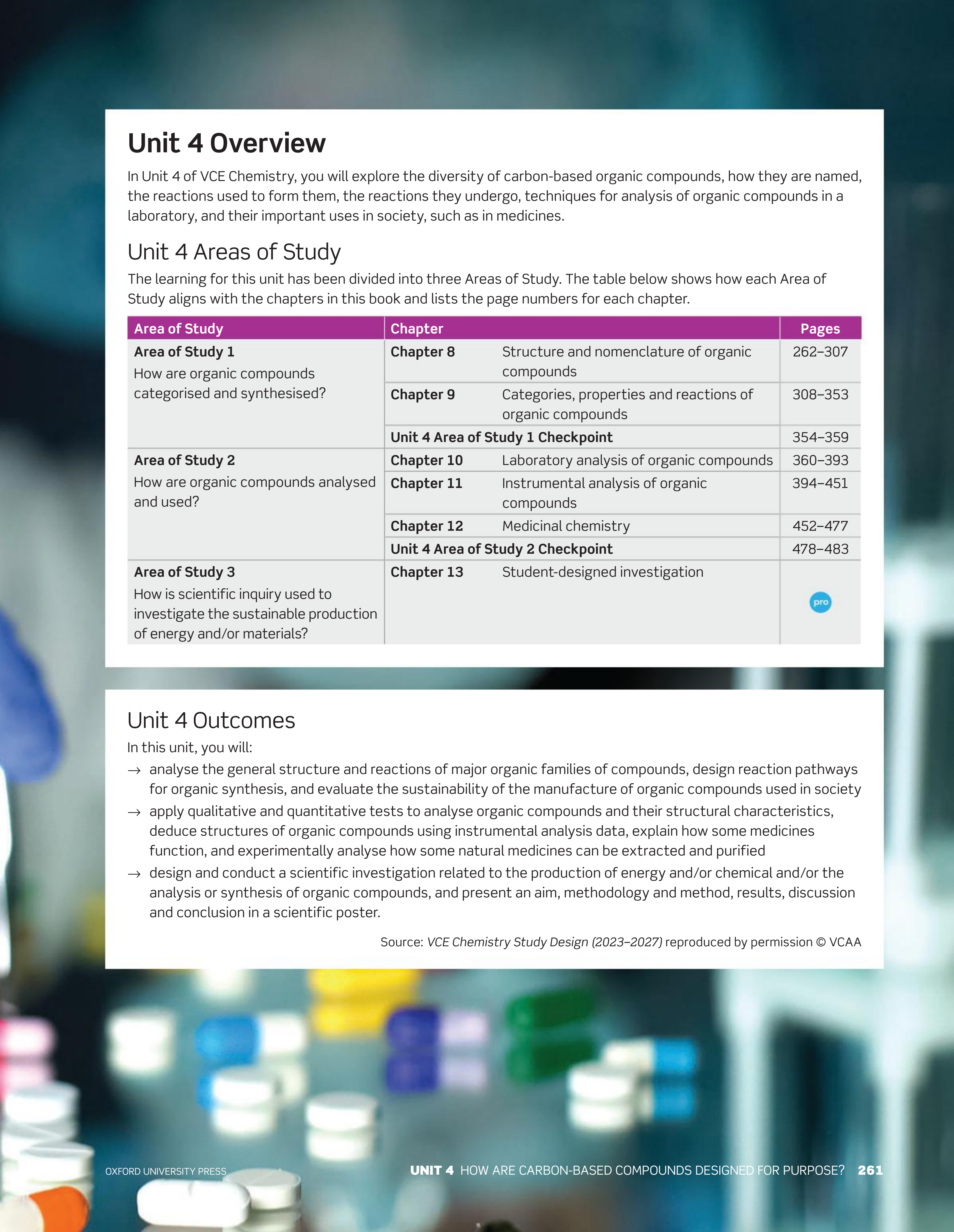
You can find the following resources for this section in your obook pro:

Unit 3 Review Part A Test your understanding questions	**Unit 3 Review Part C** Exam practice questions	**Weblink** Past examinations and examiners' reports

UNIT 4

HOW ARE CARBON-BASED COMPOUNDS DESIGNED FOR PURPOSE?

FIGURE 1 Carbon-based compounds are extremely diverse and can be used for everything from production of everyday items to medicine.



Unit 4 Overview

In Unit 4 of VCE Chemistry, you will explore the diversity of carbon-based organic compounds, how they are named, the reactions used to form them, the reactions they undergo, techniques for analysis of organic compounds in a laboratory, and their important uses in society, such as in medicines.

Unit 4 Areas of Study

The learning for this unit has been divided into three Areas of Study. The table below shows how each Area of Study aligns with the chapters in this book and lists the page numbers for each chapter.

Area of Study	Chapter		Pages
Area of Study 1 How are organic compounds categorised and synthesised?	**Chapter 8**	Structure and nomenclature of organic compounds	262–307
	Chapter 9	Categories, properties and reactions of organic compounds	308–353
	Unit 4 Area of Study 1 Checkpoint		354–359
Area of Study 2 How are organic compounds analysed and used?	**Chapter 10**	Laboratory analysis of organic compounds	360–393
	Chapter 11	Instrumental analysis of organic compounds	394–451
	Chapter 12	Medicinal chemistry	452–477
	Unit 4 Area of Study 2 Checkpoint		478–483
Area of Study 3 How is scientific inquiry used to investigate the sustainable production of energy and/or materials?	**Chapter 13**	Student-designed investigation	pro

Unit 4 Outcomes

In this unit, you will:

- analyse the general structure and reactions of major organic families of compounds, design reaction pathways for organic synthesis, and evaluate the sustainability of the manufacture of organic compounds used in society
- apply qualitative and quantitative tests to analyse organic compounds and their structural characteristics, deduce structures of organic compounds using instrumental analysis data, explain how some medicines function, and experimentally analyse how some natural medicines can be extracted and purified
- design and conduct a scientific investigation related to the production of energy and/or chemical and/or the analysis or synthesis of organic compounds, and present an aim, methodology and method, results, discussion and conclusion in a scientific poster.

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CHAPTER 8

Structure and nomenclature of organic compounds

KEY KNOWLEDGE

- characteristics of the carbon atom that contribute to the diversity of organic compounds formed, with reference to valence electron number, relative bond strength, relative stability of carbon bonds with other elements, degree of unsaturation, and the formation of structural isomers
- molecular, structural, semi-structural (condensed) formulas and skeletal structures of alkanes (including cyclohexane), alkenes, benzene, haloalkanes, primary amines, primary amides, alcohols (primary, secondary and tertiary), aldehydes, ketones, carboxylic acids and non-branched esters
- the International Union of Pure and Applied Chemistry (IUPAC) systematic naming of organic compounds up to C8, with no more than two functional groups for a molecule, limited to non-cyclic hydrocarbons, haloalkanes, primary amines, alcohols (primary, secondary and tertiary), aldehydes, ketones, carboxylic acids and non-branched esters
- trends in physical properties within homologous series (boiling point and melting point, viscosity), with reference to structure and bonding

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FIGURE 1 Organic molecules – those that contain carbon–hydrogen bonds – are often represented using a ball-and-stick model.

GROUNDWORK

In Chapter 8, you will learn about organic compounds, why they are so diverse, and how to group, name and draw them.

This chapter will build on concepts you have already learnt in Units 1 and 2. So, before you begin the chapter, test yourself by answering the following questions to make sure you remember the basics.

8A Describe covalent bonding, with reference to valence electrons.

8A Groundwork resource
Covalent bonding

8B Butane is represented by the molecular formula C_4H_{10}. Recall two other ways it can be represented.

8B Groundwork resource
Molecular representations

8C Recall the IUPAC naming rules.

8C Groundwork resource
IUPAC naming rules

PRACTICALS

8.1 **PRACTICAL: CLASSIFICATION & IDENTIFICATION**	How can the physical properties of organic compounds be used to classify them?	Page 514
8.4 **PRACTICAL: MODELLING**	How do the shapes of molecules affect their intermolecular bonding and physical properties?	pro

8.1

Carbon and organic compounds

KEY IDEAS

In this topic, you will learn that:

- carbon can bond with many other atoms because it has four valence electrons
- bond strength, bond energy and the stability of bonds all have relationships that depend on the size of the atoms involved in the bonding
- the degree of unsaturation determines the number of rings or double bonds in a molecule
- structural isomers refer to two or more organic compounds that have the same molecular formula, but a different arrangement of atoms.

organic compound
a chemical compound that contains carbon–hydrogen bonds

hydrocarbon
a compound of carbon and hydrogen atoms

Organic compounds are chemical compounds that have carbon–hydrogen bonds. There are millions of organic compounds that we know about. Compounds like methane (CH_4), glucose ($C_6H_{12}O$) and deoxyribose ($C_5H_{10}O_4$), an important part of DNA, are all organic compounds that are important to life on Earth. **Hydrocarbons** are the simplest form of organic compounds, made up of only hydrogen and carbon atoms.

Carbon plays an essential role in the diversity of organic compounds. In order to understand organic chemistry, we need to understand the characteristics of carbon atoms.

Carbon structure and bonding

valence electron
electron in the outermost shell of an atom

Carbon can form more compounds than any other element because of its structural properties. Carbon has two electrons in its inner shell and four **valence electrons** in its outer shell (Figure 1).

Study tip

Make sure you understand valence electrons and covalent bonds before you go further. You can revise these concepts in *Chemistry for VCE Units 1 and 2* Chapters 2 and 3, or you can use the groundwork resources for this chapter in your obook pro.

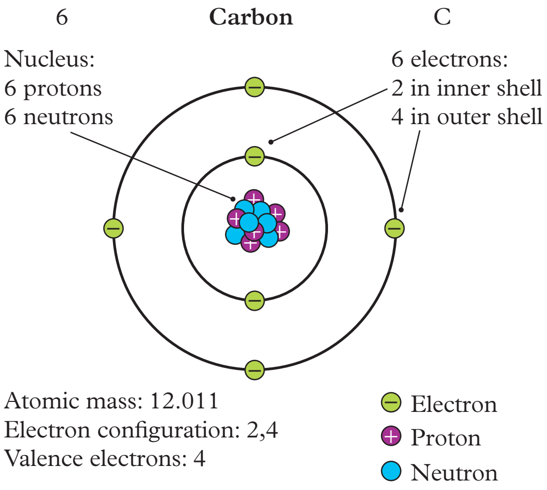


FIGURE 1 The carbon atom has four valence (outer shell) electrons.

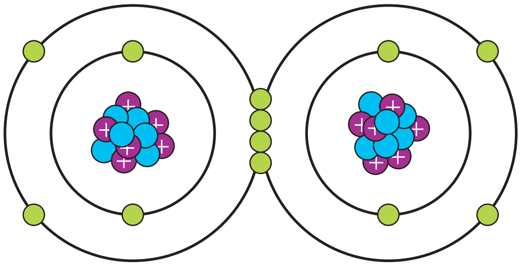


FIGURE 2 Two carbon molecules sharing two valence electrons each (covalent bonding) to form a carbon–carbon double bond.

covalent bond
a bond formed when atoms share electrons; can only occur between two non-metals

This structure allows it to:

- form single, double or triple **covalent bonds** between carbon atoms (Figure 2)
- form bonds with up to four carbon atoms
- bond with other carbon atoms to produce long carbon chains or rings (Figure 3)
- form up to four covalent bonds with non-metal atoms – for example hydrogen, oxygen, nitrogen or a halogen.

8.1
PRACTICAL: CLASSIFICATION & IDENTIFICATION

How can the physical properties of organic compounds be used to classify them? Go to page **514**.

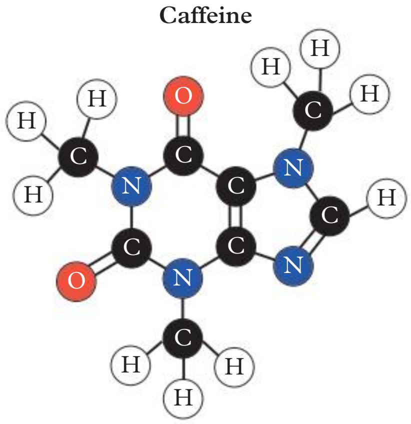


FIGURE 3 Carbon can bond with other carbon atoms to form molecules made up of carbon chains and rings – such as caffeine.

Molecular shape

Carbon's molecular shape also explains its ability to form so many compounds. In Units 1 and 2, you have learnt about the **valence shell electron pair repulsion (VSEPR) theory**, which states:

- valence electrons are grouped together (often in pairs)
- electron pairs are arranged around the central atom(s) of the molecule
- all electron pairs repel each other
- electron pairs position themselves as far apart from each other as possible.

valence shell electron pair repulsion (VSEPR) theory the theory that explains how molecules are shaped due to electrons repelling each other

When a carbon atom forms single bonds with four other atoms, the pairs of electrons in each bond act as a cloud of negative charge that repels the bonds (Figure 4).

The angle between the bonds is 109.5°. This arrangement of bonds forms a tetrahedral shape, which is the most common shape for carbon atoms (Figure 5).

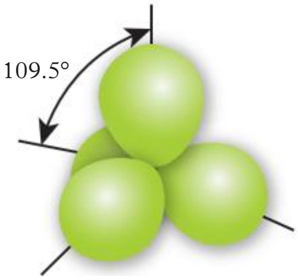


FIGURE 4 Electrons in each bond act as a cloud of negative charge that repels the bonds, placing them as far away from each other as possible.

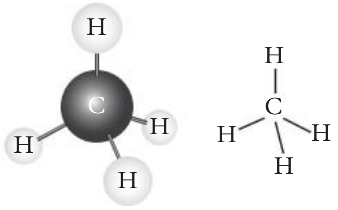


FIGURE 5 The tetrahedral shape of a carbon atom surrounded by four hydrogen atoms. All the hydrogen atoms are as far away from each other as possible.

The way carbon atoms bond together results in an endless variety of three-dimensional structures, contributing to the diversity of organic compounds.

Relative bond strength

As you learnt in Units 1 and 2, carbon forms covalent bonds. Not all covalent bonds are the same. They can differ in **bond strength**, which depends on bond energy.

Bond energy is the amount of energy required to break the covalent bonds between atoms. It is measured in kilojoules per mole (kJ mol^{-1}). There are three main factors that affect the amount of energy required to break a covalent bond, including:

- bond length
- the difference in **electronegativity** between the two atoms
- the size of the atoms.

These factors contribute to bond energy and therefore overall bond strength; you need to take all three into account when determining bond strength.

bond strength how strongly two atoms are joined together

bond energy the amount of energy (in kJ mol^{-1}) that it takes to break the covalent bond

electronegativity the ability of an atom in a bond to attract a bonding pair

TABLE 1 Common bond lengths

Bond	Average bond length (picometres)
C–H	109
C–C	154
C–N	147
C–O	143
C–F	135
C–Cl	177
C–Br	194
C–I	214
C=C	134
C=N	129
C=O	120

Bond length

Bond length is the distance between the centres of two covalently bonded molecules. The length of the bond is determined by the number of bonded electrons, the attraction between the nuclei and the size of the atoms involved. Table 1 shows some common bond lengths.

The general trend is that as you move across a period in the periodic table, the bond length decreases as you move from left to right; the bond length increases as you move down a group (Figure 9). This happens to be the same as the trend in the size of atomic radius: as the atoms get bigger, the bond length also increases, because there is more distance between the two nuclei (Figure 6).

When there are multiple bonds, the bond length is smaller than that of a single bond between the same atoms. Think of this as a rubber band attached between two spheres (Figure 7). One rubber band can be stretched quite a distance. Add a second rubber band and the distance between the two spheres gets shorter, because the resistance increases.

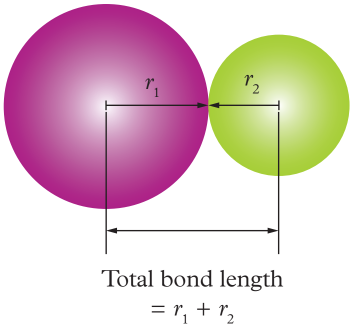


FIGURE 6 Bond length is the distance between the nuclei of the two bonded atoms.

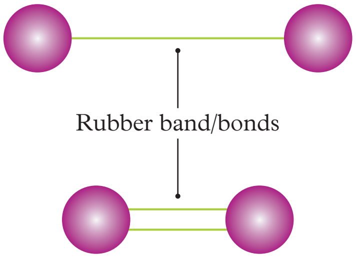


FIGURE 7 Bond length decreases as the number of bonds increases.

It is also harder to pull the two spheres apart when there are more rubber bands. This is the same with bond energy, and you will see below that the bond energy increases with more bonds between the atoms.

8.1 SKILL DRILL

Modelling bond length

Key science skill: Analyse, evaluate and communicate scientific ideas

How can you model bond length and strength using rubber bands?

Practise your skills

1 Write a step-by-step method and materials list for constructing your model.

2 Identify two potential risks of your method and how these could be made safer.

3 Use your model to describe the correlation between bond length and strength.

4 Identify possible limitations of your model. (That is, does your model represent bond length and strength accurately?)

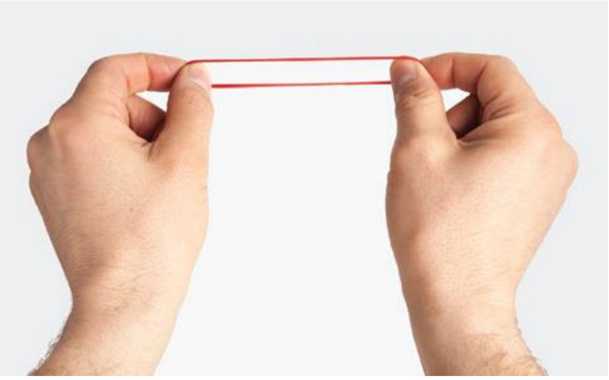


FIGURE 8 How can you use rubber bands to model bond length?

Need help analysing, evaluating and communicating scientific ideas? See Topics 1.8 and 1.10 (pages 24–27 and 30–31).

Difference in electronegativity

In covalent bonding, electrons are not always shared equally. This depends on the electronegativity of the atoms. This is an atom's ability to attract a pair of bonding electrons. As electronegativity increases, bond energy also increases. Some common values of bond energy are given in Table 2.

TABLE 2 Common bond energies

Single bonds	Energy (kJ mol^{-1} @ 273 K)	Single bonds	Energy (kJ mol^{-1} @ 273 K)	Multiple bonds	Energy (kJ mol^{-1} @ 273 K)
C–H	413	C–F	485	C=C	602
C–C	346	C–Cl	327	C=N	615
C–N	305	C–Br	285	C=O	749
C–O	358	C–I	213		

If you look at the periodic table, electronegativity increases as you move from left to right (with fluorine being the most electronegative element). Bond energy also increases (Figure 9).

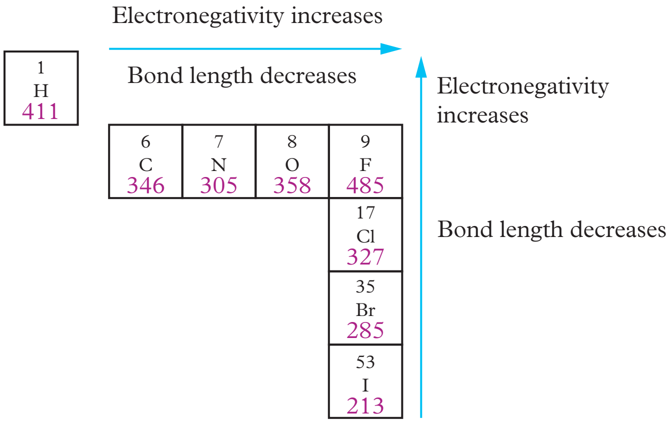


FIGURE 9 The relationship between bond strength (purple numbers) and electronegativity

Relative bond stability

The more energy is needed to break a covalent bond, the more stable the bond is. In other words, the higher the bond energy, the greater the **bond stability**. Bonds with low energy will be less stable as they require less energy to break them.

bond stability
the more energy that is required to break the bond, the more stable the bond is

You can think back to the rubber bands and the spheres in Figure 7 – it will take more energy to pull apart the two spheres attached by two rubber bands than the spheres with the single rubber band. So, this means that double bonds are more stable than single bonds between the same two atoms.

FIGURE 10 The more stable a bond, the more heat energy that must be applied to break it.

Degree of unsaturation

saturated (molecule)
a molecule that contains only single carbon bonds (C–C)

unsaturated (molecule)
a molecule that contains at least one double or triple bond

cyclic
a compound that contains atoms connected in a ring

non-cyclic
a compound that contains atoms connected only in single or branched chains

A molecule is considered **saturated** if it only contains single bonds in its structure. Saturated molecules have the maximum number of hydrogen atoms attached to all carbon atoms. A molecule is **unsaturated** if it contains one or more double or triple bonds. Unsaturated molecules do not have the maximum number of hydrogen atoms attached to all carbon atoms.

Consider the molecules propane and propene in Figure 11. Propane is saturated, meaning that each carbon atom has all single bonds and is attached to the maximum number of hydrogen atoms. Propene is unsaturated, meaning it has a double bond and does not have the maximum number of hydrogen atoms attached to all the carbon atoms.

An exception to this rule is **cyclic** alkanes that do not contain double or triple bonds, such as cyclohexane (C_6H_{12}) in Figure 12. They do not have the maximum number of hydrogens, but they are considered saturated. Therefore, the rule only applies to **non-cyclic** compounds.

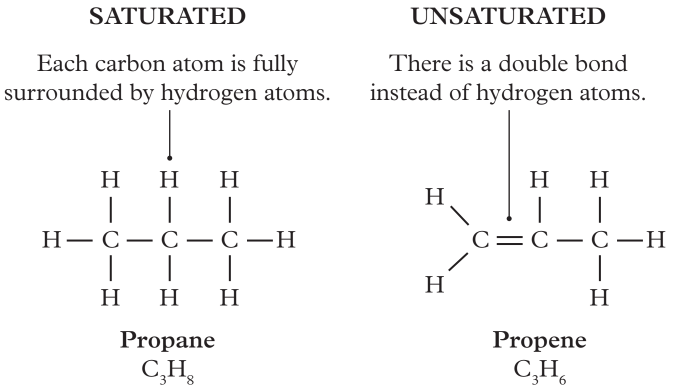


FIGURE 11 Propane and propene

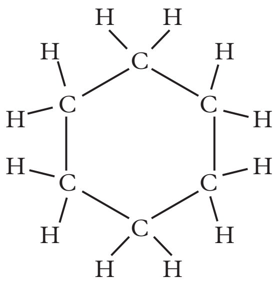


FIGURE 12 Cyclohexane

degree of unsaturation
the number of double bonds or rings in the structure.

The maximum number of hydrogen atoms that can be attached to a carbon atom is represented by $2C + 2$. Each double bond or ring reduces the number of hydrogen atoms in the molecule by 2. The number of double bonds or rings is known as the **degree of unsaturation** (Table 3 on the next page).

The degree of unsaturation (DoU) can be calculated using the formula:

$$\text{DoU} = \frac{2C + 2 + N - X - H}{2}$$

where:

C = number of carbon atoms

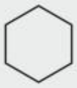
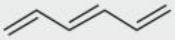
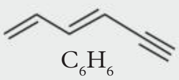
N = number of nitrogen atoms

X = number of halogens (Cl, F, Br, I)

H = number of hydrogen atoms

Nitrogen and halogens are accounted for in this equation. For a molecule with nitrogen to be fully saturated, it must contain one more hydrogen molecule, as nitrogen will attach to a hydrocarbon as NH_2. Also, if a halogen is attached to a hydrocarbon, it will take the place of a hydrogen molecule, so there will be one less hydrogen atom attached. Oxygen is not accounted for in the equation, because saturation is unaffected by oxygen. For example, propanol (C_3H_7OH) has the same number of hydrogen atoms as propane (C_3H_8).

TABLE 3 Degrees of unsaturation

Degree of unsaturation	Bonds or rings	Examples
0	Saturated: no rings or double bonds (C=C)	 C_6H_{14}
1	1 C=C or 1 ring	 C_6H_{12}  C_6H_{12}
2	2 C=C or 2 rings or 1 C=C and 1 ring	 C_6H_{10}  C_6H_{10}
3	3 C=C or 3 rings or 2 C=C and 1 ring or 2 rings and 1 C=C	 C_6H_8  C_6H_8
4	4 C=C or 4 rings or 3 C=C and 1 ring or any combination of the 2, equalling 4	 C_6H_6  C_6H_6

Study tip

Although cyclohexane does not have the maximum number of hydrogens and has a DoU of 1, it is considered saturated as it does not contain double or triple bonds.

8.1A Worked example
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8.1A Worked example
Video demonstration

In Worked example 8.1A, you can see this formula being used to calculate the DoU for different organic compounds.

Structural isomers

Isomers are molecules that contain the same number and type of atoms, but the atoms are arranged in different ways. Even though they might have the same atoms, their different arrangement can create unique molecules with their own physical and chemical properties. The presence of isomers is another reason for the diversity of organic compounds.

Isomers can be either **structural isomers** or stereoisomers. You will learn about stereoisomers in Chapter 12.

Structural isomers can be **chain isomers**, **positional isomers**, or a combination of both.

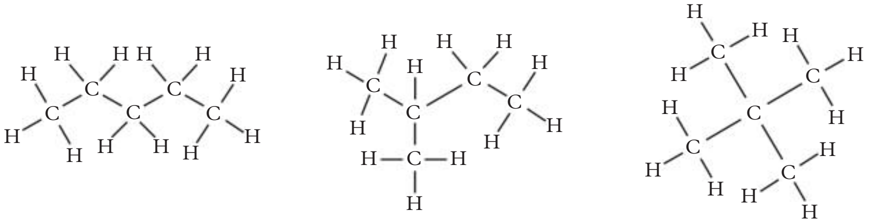
structural isomer
a molecule with the same molecular formula as another molecule, but different structure

chain isomer
a structural isomer caused by branching of the hydrocarbon chain

positional isomer
a structural isomer where the functional group is in different positions on the same carbon chain

Chain isomers

Chain isomers occur when molecules have the same molecular formula, but a different arrangement or branching of the main carbon chain (Figure 13).



Pentane **2-methylbutane** **2,2-dimethylpropane**

FIGURE 13 Three of the chain isomers of C_5H_{12}: pentane, 2-methylbutane and 2,2-dimethylpropane

alkyl chain
the general name for a main hydrocarbon chain consisting of any number of carbon atoms

In pentane, the main chain or **alkyl chain** has five carbon atoms, as shown in Figure 13 on the previous page. In 2-methylbutane, a methyl group (CH_3) has replaced a hydrogen in the main chain carbon. In other molecules, this could be an ethyl (CH_2CH_3) or propyl ($CH_2CH_2CH_3$) group. These hydrocarbon branches are called alkyl groups. You might remember this from Units 1 and 2, but we will talk about them again in Topic 8.3.

Positional isomers

main chain
also known as the parent chain, parent molecule or longest chain

Positional isomers occur when molecules have the same molecular formula, but a double bond or functional group is attached to a different position along the **main chain**.

8.1B Worked example
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8.1B Worked example
Video demonstration

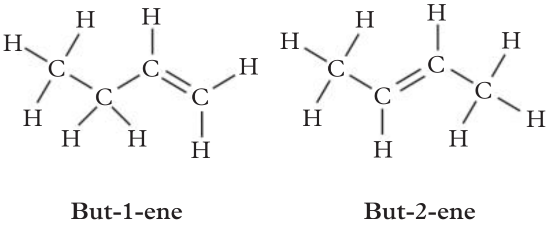


FIGURE 14 Positional isomers of butene C_4H_8: but-1-ene and but-2-ene

In the positional isomers of butene, only the position of the double bond changes. The main chain stays the same (there are no branches). Positional isomers also exist where the locations of other functional groups along the main chain are different.

In Worked example 8.1B, you will learn how to identify structural isomers.

8.1 CHECK YOUR LEARNING

Describe and explain

1 Define the term *organic compound*.

2 Describe two characteristics of the carbon atom that contribute to the diversity of organic compounds.

3 Explain how bond energy is related to bond strength.

4 Distinguish between saturated and unsaturated molecules.

5 Explain how structural isomers contribute to the diversity of organic compounds.

6 Distinguish between chain and positional isomers.

Apply, analyse and compare

7 Analyse the following and identify if this is the correct order of bond length:
C=N > C–Br > C–O > C–C

8 Apply your understanding of bond strength by putting these in the correct order of bond strength (highest to lowest):
C=N, C=O, C–C, C–H, C–F

9 Calculate the degree of unsaturation for the following molecules:
a $C_7H_{15}F$ **b** C_4H_7OH **c** C_6H_{12}

10 Consider the molecular formula C_6H_{12}. Determine the number of non-cyclic isomers and draw them all.

Design and discuss

11 Discuss the relationship between the atomic radius and bond length.

FIGURE 15 The position of a double bond can change the state of a substance. For example, butter is solid at room temperature and oil is a liquid.

8.2

Functional groups

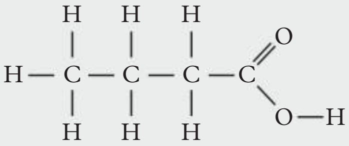
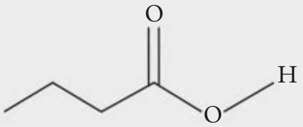
KEY IDEAS

In this topic, you will learn that:

- there are four ways to represent organic compounds: molecular formulas, structural formulas, semi-structural formulas and skeletal structures
- functional groups are groups of atoms that change the properties of organic molecules.

There are four ways to represent organic compounds. They are shown in Table 1, using butanoic acid as an example. Worked example 8.2 shows you how to convert between structural representations.

TABLE 1 Four ways to represent butanoic acid as an organic compound

Formula	Representation	Detail
Molecular formula	$C_4H_8O_2$	Gives no information about the structure of a molecule; best used when writing chemical equations.
Structural formula		Gives the exact layout of the structure of the molecule; shows *all* the atoms and how they are bonded together.
Semi-structural or condensed formula	$CH_3CH_2CH_2COOH$ or $CH_3(CH_2)_2COOH$	Shows the position of all the atoms and functional groups. In the second, condensed form, the consecutive CH_2 segments are grouped with brackets and a subscript for the number of CH_2 segments in a row. You can also show C=C if they appear in the molecule.
Skeletal structure		Does not show carbon or hydrogen atoms but instead, uses vertices (corners) to represent the carbon atoms (with hydrogen atoms attached); shows all functional groups.

8.2 Worked example
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8.2 Worked example
Video demonstration

Homologous series

A **homologous series** is a group of hydrocarbons that have a similar structure, similar chemical properties, and a trend in physical properties (explored in Topic 8.4). Each group is represented by the same general formula.

homologous series
a sequence of compounds that share the same characteristics; also known as hydrocarbon families

Alkanes

Alkanes are saturated hydrocarbons because they only contain single carbon bonds. Alkanes can consist of a single chain of hydrocarbons (or alkyl chain), or exist as structural isomers.

Alkanes have the general formula C_nH_{2n+2}, where n is the number of carbon atoms in the molecule. For example, if an alkane has three carbon atoms, we can substitute this number in the formula to find the number of hydrogen atoms: $C_3H_{(2\times3+2)} = C_3H_8$.

Many alkanes are used as fuels. Natural gas is made from methane, petrol is made from octane and LPG is made from propane (Figure 1).

alkane
a saturated hydrocarbon that only contains single covalent bonds

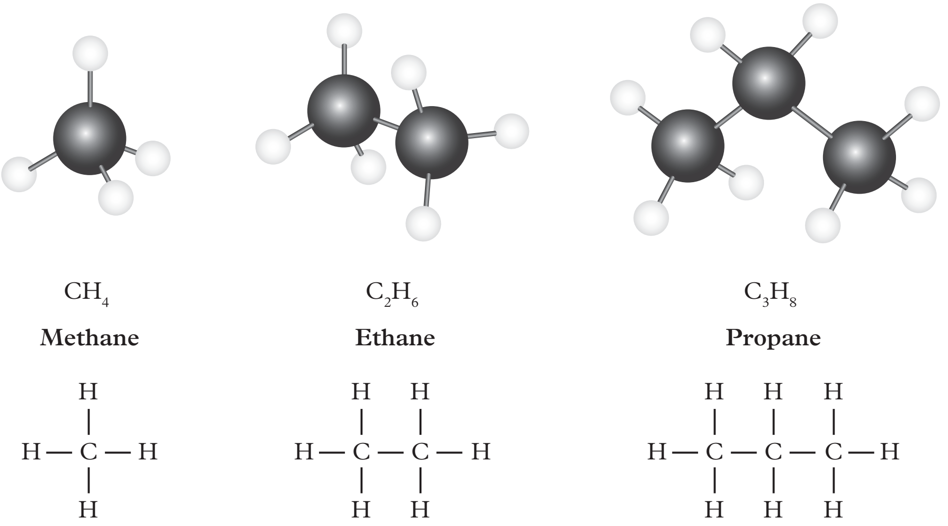


FIGURE 1 The first three alkane structures: methane, ethane and propane

Naming alkanes

suffix
a letter or group of letters added to the end of a word to change its meaning or create a new word

Alkanes have the **suffix** (the end of the word) '-ane' (Figure 2). The names for alkanes depend on the number of carbon atoms present in the molecule.

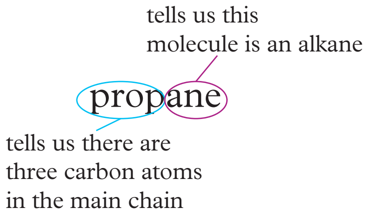


FIGURE 2 Alkanes use the suffix '-ane' in their name.

prefix
a letter or group of letters added to the beginning of a word to change its meaning or create a new word

You can name alkanes using the following steps.

1. Number the carbon atoms in the longest carbon chain.
2. Use the **prefix** for the number of carbon atoms present (e.g. if there is one carbon atom, choose 'meth-'; for two carbon atoms, choose 'eth-', and so on). These are listed in Table 2.
3. Use the suffix '-ane' to show there are only single covalent bonds in the molecule.

Study tip

You might remember the IUPAC naming system for organic compounds from Units 1 and 2. You will revise this and learn about how to name compounds with new functional groups in the next topic.

TABLE 2 Naming alkanes

Number of carbon atoms	Molecular formula	Prefix	Name
1	CH_4	Meth-	Methane
2	C_2H_6	Eth-	Ethane
3	C_3H_8	Prop-	Propane
4	C_4H_{10}	But-	Butane
5	C_5H_{12}	Pent-	Pentane
6	C_6H_{14}	Hex-	Hexane
7	C_7H_{16}	Hept-	Heptane
8	C_8H_{18}	Oct-	Octane

Cyclohexane

Alkanes are mostly arranged in linear structures; however, they can also be arranged in a cyclic or ring structure, where the alkyl chain closes. These are called **cycloalkanes**. The prefix ‘cyclo-’ tells us that a molecule has a ring structure (Figure 3).

cycloalkane
a hydrocarbon compound containing only single C–C bonds, in which the carbon chain closes to form a ring

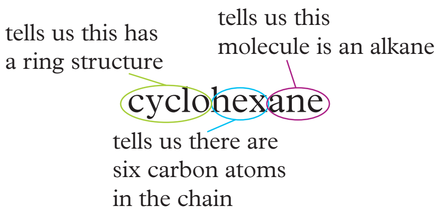


FIGURE 3 Cyclic structures such as cyclohexane begin with the prefix ‘cyclo-’.

Cyclohexane (C_6H_{12}) is an example of a cycloalkane. It is a cyclic isomer of C_6H_{12}. As seen in Figure 4, all the carbon atoms are bonded in a ring, so it does not have any **terminal carbon** atoms. Each carbon atom still has four bonds, and in this structure, they are all CH_2.

Although it is called cyclohex*ane*, it does not have the same general formula as the alkanes. This is because it has no terminal carbon atoms. It has the same general formula as the alkenes, C_nH_{2n}.

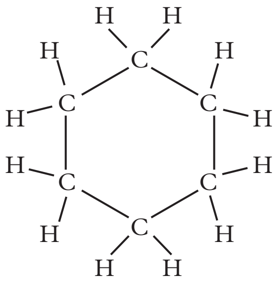


FIGURE 4 Structure of cyclohexane

terminal carbon
a carbon atom at the end of a hydrocarbon chain

Functional groups

Compounds in a homologous series can also contain **functional groups**, which are atoms or groups of atoms that give specific chemical properties to the molecule they are part of. While the hydrocarbons provide the backbone for a lot of organic compounds, functional groups define their behaviour and function.

functional group
specific group of distinctly arranged atoms that changes the properties of a molecule

The functional groups you need to learn for Unit 4 include double bonds (in alkenes), haloalkanes, primary amines and amides, alcohols, aldehydes, ketones, carboxylic acids and esters.

Alkenes

Alkenes are unsaturated hydrocarbons, because they contain a carbon–carbon double bond as well as carbon–carbon single bonds. These carbon–carbon double bonds are considered functional groups because they give the organic compound special chemical properties. The double bond replaces a hydrogen in the alkyl chain or group.

They have the general formula C_nH_{2n}, where n is the number of carbon atoms in the molecule. For example, if an alkene has five carbon atoms, we can substitute this number in the formula to find the number of hydrogen atoms: $C_5H_{(2\times5)} = C_5H_{10}$.

Naming alkenes

Alkenes end in ‘-ene’. Alkenes use the same prefixes as alkanes, which depend on the number of carbon atoms present in the molecule.

FIGURE 5 Alkenes are often used to make products like the acrylic fibres in yarn.

You can name alkenes using the following steps (Figure 6).

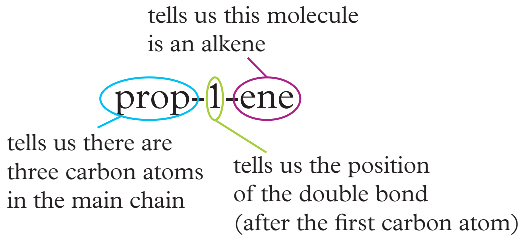
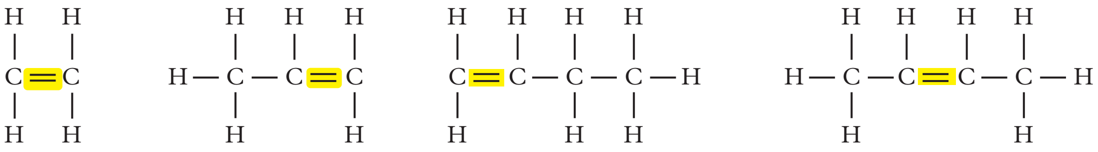


FIGURE 6 Alkene names use the suffix '-ene' and indicate the position of the double bond.

1. Number the carbon atoms in the chain, starting at the end closest to the double bond.
2. Identify the position of the double bond by finding the number of the first carbon atom in the double bond.
3. Insert that number in front of the ending '-ene'.

In Figure 7 you'll notice that the first alkene is named ethene. There is no number in this molecule's name. This is because there is only one place for the double bond to be positioned (between the only two carbon atoms). We therefore write ethene because it is unnecessary to write 1-ethene.

The fourth alkene is named but-2-ene, rather than butene. This is because in the molecule C_4H_8, the double bond could be in two places: after the first carbon atom, or after the second. In order to correctly identify the molecule, you need to specify its name as but-1-ene or but-2-ene.



Ethene **Prop-1-ene** **But-1-ene** **But-2-ene**

FIGURE 7 Molecular and structural formulas for the first four alkenes

Study tip

There is no such thing as 'methene'. If a molecule only has one carbon atom, there is no place for a double bond to go.

The placement of the double bond is important for IUPAC naming, which we will discuss more in Topic 8.3.

Benzene

Benzene, C_6H_6, is not an alkene, although the suffix '-ene' can be confusing. It is an unsaturated cyclic hydrocarbon made up of six carbon atoms, with three carbon–carbon double bonds in a ring structure. Each carbon atom in a benzene molecule is attached to a single hydrogen atom.

Benzene is often shown in diagrams with alternating double and single bonds (Figure 8), but this is not entirely correct, as the electrons are not shared exclusively by two carbon atoms all the time. The electrons are actually **delocalised** – meaning they are free to move around the benzene ring structure. For this reason, benzene can be drawn as in Figure 9.

delocalised
(when referring to electrons) not associated with a single atom or bond and therefore free to move around the chemical structure

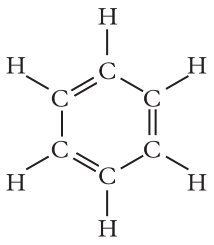


FIGURE 8 A common representation of benzene

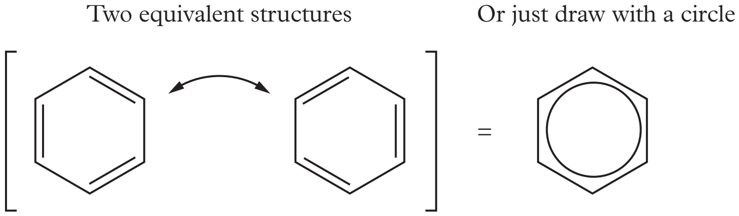


FIGURE 9 Benzene can be drawn with alternating double bonds, or with a circle to indicate the delocalised electrons.

Benzene has a similar cyclic structure to cyclohexane, but differs in that it has a ring of delocalised electrons and one less hydrogen per carbon atom.

Benzene is classed as an **aromatic hydrocarbon** and is found naturally in crude oil. It is used as an intermediate to make other chemicals such as types of polystyrene, rubbers, lubricants, dyes, detergents, drugs, explosives, and pesticides.

aromatic hydrocarbon
a hydrocarbon compound that contains a ring with delocalised electrons

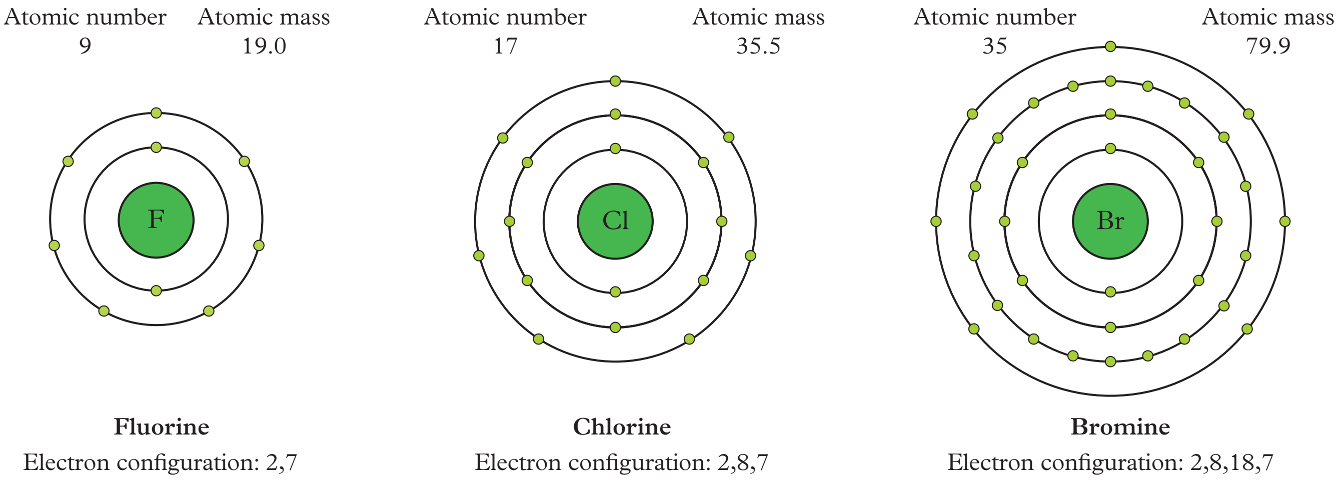
Haloalkanes

Haloalkanes are organic compounds that contain one or more halogen. Halogens are the elements of group 17 on the periodic table: fluorine (F), chlorine (Cl), bromine (Br), iodine (I), astatine (At), and tennessine (Ts). In a haloalkane, each halogen replaces a hydrogen in the alkyl chain or group. The general formula for haloalkanes is R–X, where R is the alkyl chain or group and X is the halogen.

Each of the halogen atoms has seven valence electrons (Figure 10). This means they can share their single unpaired electron with one of the four unpaired valence electrons in a carbon atom to form a single covalent bond.

As the halogens are more electronegative than carbon, they will create a polar bond with carbon. This results in the halogens attracting the shared pair closer to their own nucleus, so the carbon atom will have a slight positive charge (or **dipole**) and the halogen will have a slight negative charge or dipole.

dipole
unequal sharing of electrons creating a slight negative or positive charge at each end of the bond



Fluorine
Electron configuration: 2,7

Chlorine
Electron configuration: 2,8,7

Bromine
Electron configuration: 2,8,18,7

FIGURE 10 The halogen electron shells

Naming haloalkanes

Haloalkanes use the same main chain suffix as the alkanes, as shown in Figure 12. You can name haloalkanes using the following steps (Figure 12).

$$\overset{\delta^+}{H_3C} — \overset{\delta^-}{Cl}$$

FIGURE 11 The carbon–chlorine bond creates dipoles for the carbon (slight positive charge) and chlorine (slight negative charge) atoms.

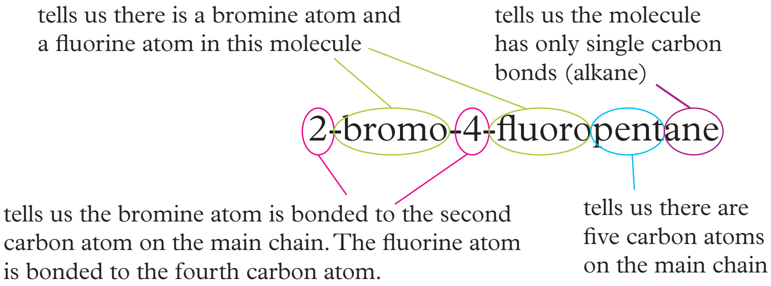


FIGURE 12 Halogen atoms are named ending in '-o' when they are part of a haloalkane.

1 Number the carbon atoms in the longest chain containing halogen(s).

2 Insert the prefix for the halogen(s) (Table 3) at the beginning of the molecule's name and '-ane' to the end.

3 If there is more than one of the same halogen present, use a prefix for the number (e.g. 'di-' for two, 'tri-' for three etc.)

4 If more than one halogen is present, list them in alphabetical order.

5 Identify the position of the halogen(s) by finding the number of the main-chain carbon atom it is bonded with.

6 Insert that position number in front of the halogen functional group.

Fluoroethane (C_2H_5F) is an example of a haloalkane (Figure 13). One of the hydrogen atoms from the alkane has been replaced by the halogen.

TABLE 3 Names of haloalkane groups

Halogen	Prefix
Fluorine	Fluoro-
Chlorine	Chloro-
Bromine	Bromo-
Iodine	Iodo-

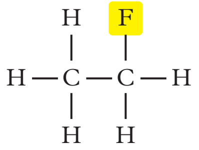


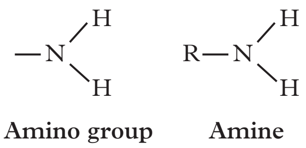
FIGURE 13 Fluoroethane

Primary amines

Amines contain an **amino functional group**. This consists of a nitrogen atom with two hydrogen atoms attached to it (NH_2). The nitrogen then forms a bond with a carbon atom in an alkane. Therefore, each amino group replaces a hydrogen in the alkyl chain or group.

A **primary amine** is always an NH_2 with two unbonded electrons, and it occurs when the amino group is bonded to one carbon. The general formula for primary amines is $R–NH_2$, where R is an alkyl chain or group.

You do not need to know about secondary or tertiary amides in VCE Chemistry, so they will not be covered here.



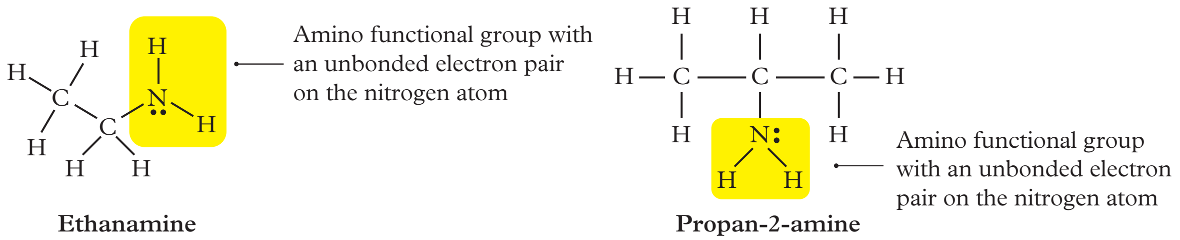
Amino group **Amine**

FIGURE 14 An amino group is an NH_2 molecule that will attach to a carbon atom to form an amine. R represents an alkyl chain or group.

amine
a homologous series that contains the functional group NH_2

amino functional group
an NH_2 molecule that will attach to a carbon atom to form an amine

primary amine
an amine where nitrogen is always bonded to one carbon and two hydrogen atoms



Amino functional group with an unbonded electron pair on the nitrogen atom

Amino functional group with an unbonded electron pair on the nitrogen atom

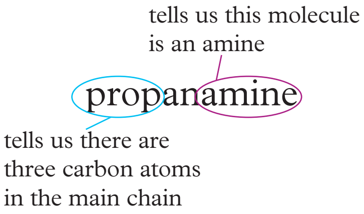
Ethanamine **Propan-2-amine**

FIGURE 15 The structures of ethanamine and propan-2-amine are both examples of primary amines

Naming primary amines

Primary amines follow the same naming structure as all hydrocarbons. The same names are used for the main chain as for the alkanes, and '-amine' forms the suffix of the name.

You can name primary amines using the following steps (Figure 16).



tells us this molecule is an amine

tells us there are three carbon atoms in the main chain

FIGURE 16 Propanamine is a primary amine with three carbon atoms on the main carbon chain.

1 Number the carbon atoms in the longest carbon chain.
2 Use the prefix for the number of carbon atoms present (e.g. if there are two carbon atoms, choose ‘eth-’; for three carbon atoms, choose ‘prop-’, and so on).
3 Add the alkane suffix ‘-ane’ but replace ‘-e’ with ‘-amine’ at the end of the name to show there is an NH_2 present.

Sometimes, an amine can also be added as the prefix ‘amino’ before the main chain name. This depends on what other functional groups are present. We will discuss this in Topic 8.3, when we look at priority naming.

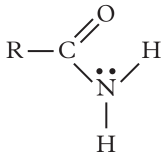


FIGURE 17 Primary amide: a carbon atom with a double-bonded oxygen atom and an NH_2 group attached to the alkyl chain or group R

Primary amides

Like amines, amides also contain an NH_2. However, a **primary amide** also has a double-bonded oxygen atom attached to the carbon atom that the NH_2 is attached to (Figure 17).

The general formula for a primary amide is R–$CONH_2$. Primary amides are always on the end of a hydrocarbon chain. In Figure 18, this is represented by R (R is an alkyl chain or group). You do not need to know about secondary or tertiary amides in VCE Chemistry, so they will not be covered here.

primary amide contains a carbon atom with a double-bonded oxygen atom and NH_2 group attached to the alkyl chain or group R

Study tip

To remember the difference between amines and amides, think about the ‘d’ in amide – it can stand for double bond. Amine has no ‘d’ in the name and there is no double bond.

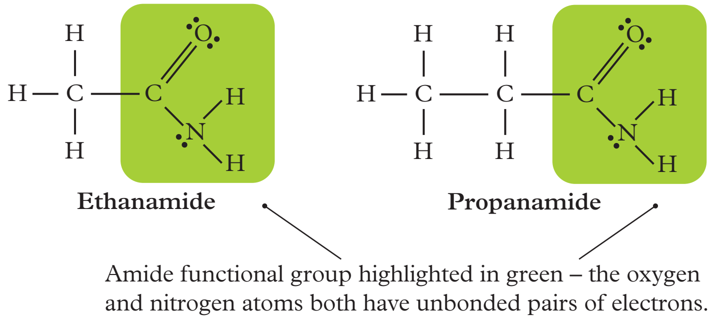


FIGURE 18 Ethanamide and propanamide are both primary amides, as the amide functional group is on the end of the chain.

Naming primary amides

Primary amides also use the same names for the main chain as the alkanes, while the amide forms the suffix of the name.

You can name primary amides using the following steps (Figure 19).

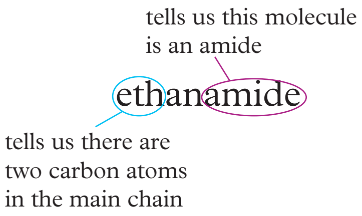


FIGURE 19 The suffix ‘-amide’ is used to name something as a primary amide. Ethanamide is a primary amide with two carbon atoms on the main carbon chain.

1 Number the carbon atoms in the longest carbon chain.
2 Use the prefix for the number of carbon atoms present (e.g. if there are two carbon atoms, use ‘eth-’; for three carbon atoms, choose ‘prop-’, and so on).
3 Add the alkane suffix ‘-ane’, but replace ‘-e’ with ‘-amide’ at the end of the name to show there is a $CONH_2$ present.

Alcohols

Alcohols are hydrocarbons that have a **hydroxyl group** attached to them. A hydroxyl group is an oxygen and a hydrogen atom (–OH), shown in Figure 20. Each hydroxyl group replaces a hydrogen in the alkyl chain or group. The general formula for alcohols is R–OH.

$R - O - H$

FIGURE 20 A hydroxyl group is an oxygen and a hydrogen atom (–OH) attached to an alkyl chain or group R.

hydroxyl group an oxygen and a hydrogen atom (–OH) attached to an alkyl chain or group R

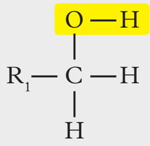
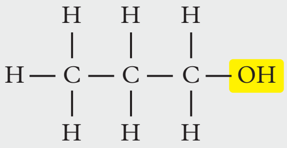
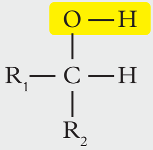
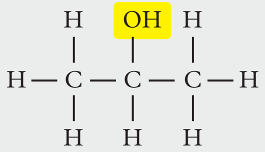
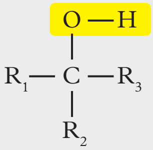
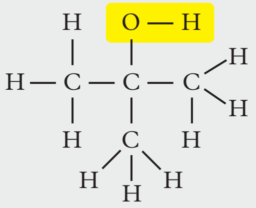
Study tip

Hydroxyl is not the same as hydroxide. Hydroxyl is a covalent –OH group bonded to a carbon chain in an organic molecule. Hydroxide is a negatively charged OH ion found in ionic substances or dissolved in water.

Primary, secondary and tertiary alcohols

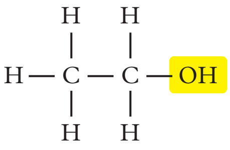
Alcohols can be classified as primary, secondary or tertiary. This depends on how many alkyl groups are attached to the carbon atom that is bonded to the hydroxyl group (Table 4).

TABLE 4 Features of primary, secondary and tertiary alcohols

Type of alcohol	Features	General formula	Example
Primary (1°) alcohol	• The carbon atom bonded to the –OH is attached to one other carbon atom or alkyl group (R_1). • The –OH is always on the terminal carbon atom in a hydrocarbon chain.		 **Propan-1-ol**
Secondary (2°) alcohol	• The carbon atom bonded to the –OH is attached to two other carbon atoms or alkyl groups (R_1 and R_2). • The –OH can be positioned at any carbon atom along the hydrocarbon chain, except a terminal carbon atom.		 **Propan-2-ol**
Tertiary (3°) alcohol	• The carbon atom bonded to the –OH is attached to three other carbon atoms or alkyl groups (R_1, R_2 and R_3). • The hydroxyl group is attached to a carbon atom that has no hydrogen atoms attached to it – only alkyl groups.		 **2-methylpropan-2-ol**

Study tip

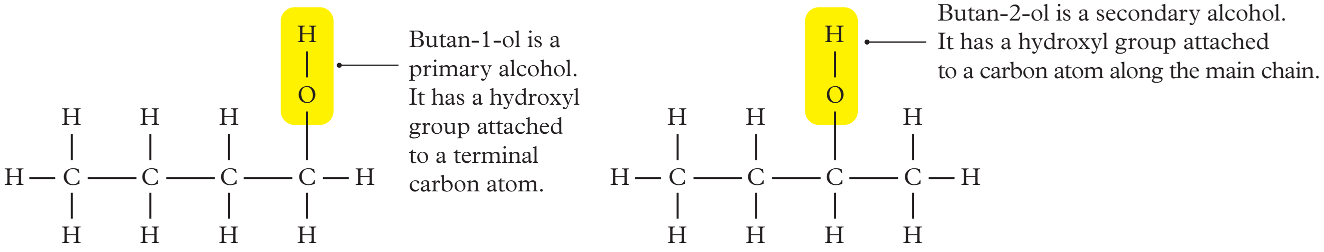
The placement of the hydroxyl group will be important in Topic 8.4 when we discuss boiling and melting points of alcohols.



Ethanol

FIGURE 21 Ethanol is always a primary alcohol.

Figure 22 shows the difference between the placement of the hydroxyl group of primary and secondary alcohols. In butan-1-ol, the hydroxyl functional group is on the end (terminal) carbon atom; in butan-2-ol, the hydroxyl group is on a carbon in the middle of the main chain. Ethanol is always a primary alcohol as it only has two carbon atoms in its chain.



Butan-1-ol is a primary alcohol. It has a hydroxyl group attached to a terminal carbon atom.

Butan-2-ol is a secondary alcohol. It has a hydroxyl group attached to a carbon atom along the main chain.

FIGURE 22 Butan-1-ol is a primary alcohol and butan-2-ol is a secondary alcohol.

Naming alcohols

Alcohols use the same names for the main chain as other hydrocarbons, and the suffix of the name is '-ol'.

You can name alcohols using the following steps (Figure 23).

1 Number the longest chain containing the –OH group (starting at the end closest to the –OH group).
2 Number any other groups according to their position on the carbon chain.
3 Add the alkane suffix '-ane', but replace '-e' with '-ol'.

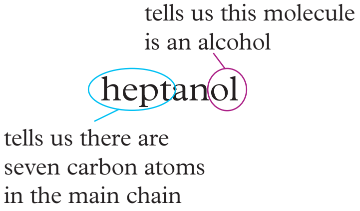


FIGURE 23 The suffix '-ol' is used to name a molecule as an alcohol (in this case heptanol).

Sometimes hydroxyl can form the prefix, becoming 'hydroxy-' if certain functional groups are attached to the molecule. We will go over this in more detail in the next topic.

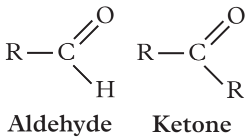


FIGURE 24 Aldehyde and ketone functional groups

Aldehydes and ketones

Aldehydes and ketones both contain a **carbonyl group**. A carbonyl group is a carbon atom double-bonded to an oxygen atom (C=O). The difference between the two is that:

- an **aldehyde** has the carbonyl group on a terminal carbon atom. It must contain at least one CHO group. The general formula for aldehydes is therefore R–COH.
- a **ketone** has the carbonyl group on a carbon atom along the main chain. The carbonyl must be attached to a carbon atom on either side. The general formula for ketones is therefore R–CO–R. Note that the alkyl chain length for the R groups can be different. This is shown in Figure 24.

carbonyl group
a carbon atom double-bonded to an oxygen atom

aldehyde
contains a carbonyl group on a terminal carbon atom

ketone
contains a carbonyl group on a carbon atom along the main chain

Naming aldehydes or ketones

Aldehydes and ketones use the same names for the main chain as other hydrocarbons. The suffix of the name is: '-al' for aldehydes and '-one' for ketones.

You can name aldehydes and ketones using the following steps (Figures 25 and 26).

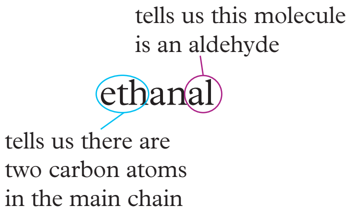


FIGURE 25 The suffix '-al' is used to name aldehydes.

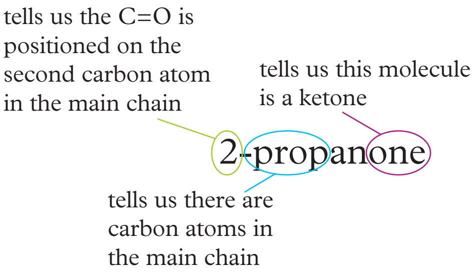
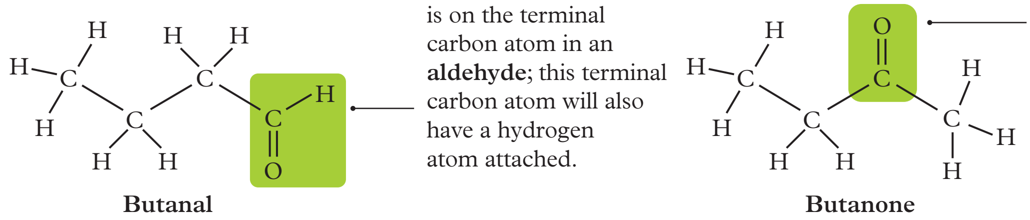


FIGURE 26 The suffix '-one' is used to name ketones.

1 Number the longest chain containing CHO group.
2 Use the prefix for the number of carbon atoms present (e.g. if there are two carbon atoms, choose 'eth-'; for three carbon atoms, choose 'prop-', and so on).
3 Add the alkane suffix '-ane', but replace '-e' with '-al' for aldehydes or '-one' for ketones.
4 For ketones, the position of C=O is represented by the position of the carbon atom on the main chain.

The carbon atom with the carbonyl group is counted as a part of the main chain, so an aldehyde with four carbon atoms is butanal and a ketone with four carbon atoms is butanone (Figure 27).



The carbonyl group is on the terminal carbon atom in an **aldehyde**; this terminal carbon atom will also have a hydrogen atom attached.

The carbonyl group is along the main chain in a **ketone**. The carbon atom with the carbonyl group will not have any hydrogen atoms attached.

FIGURE 27 Butanal and butanone structures

Carboxylic acids

carboxylic acid contains a carboxyl group

carboxyl group a hydroxyl group bonded to a carbon atom, which has a double-bonded oxygen atom

Like alcohols, **carboxylic acids** have a hydroxyl group (–OH) attached to a carbon atom. The carbon atom also has a double-bonded oxygen. This functional group is called a **carboxyl group** (Figure 28). The general formula for a carboxylic acid is R–COOH.

The COOH is always positioned at the end of a carbon chain (Figure 29). This is because the carbon uses three of its four valence electrons to bond with oxygen atoms in the carboxyl group.

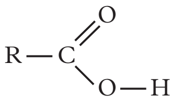


FIGURE 28 The carboxyl functional group is –COOH.

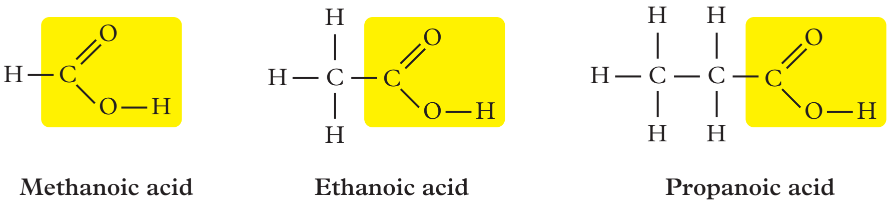
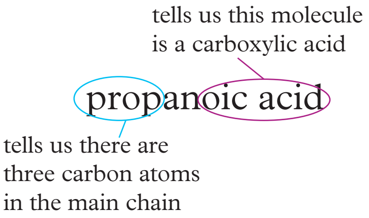


FIGURE 29 The first three carboxylic acids – the carboxyl groups are highlighted.

Naming carboxylic acids

Carboxylic acids use the same names for the main chain, with the suffix ‘-oic acid’.

You can name carboxylic acids using the following steps (Figure 30).



tells us this molecule is a carboxylic acid

tells us there are three carbon atoms in the main chain

FIGURE 30 Carboxylic acids are written with ‘-oic acid’ as a suffix.

1 Number the longest carbon chain. The carbon atom in the carboxyl group is counted as part of the main chain.

2 Use the prefix for the number of carbon atoms present (e.g. if there are two carbon atoms, choose ‘eth-’; for three carbon atoms, choose ‘prop-’, and so on).

3 Add the alkane suffix ‘-ane’, but replace ‘-e’ with ‘-oic acid’.

FIGURE 31 Do you recognise the smell of squashed ants? It’s the smell of methanoic acid – a pungent carboxylic acid.

Esters

Esters are the result of a reaction between an alcohol and a carboxylic acid. You will learn more about this reaction in Chapter 9.

The ester functional group:

- has a double-bonded oxygen atom (carbonyl group)
- has an oxygen atom that links two carbon atoms
- appears in the middle of a carbon chain (see Figure 32).

The general formula for esters is R–COO–R. The alkyl chain length for the R groups can be different.

Esters are often responsible for natural and synthetic flavours and smells, such as those found in foods and perfumes (Figure 33).

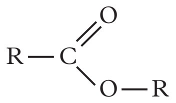
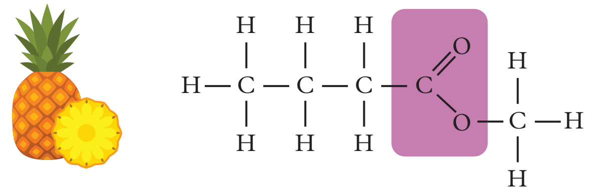
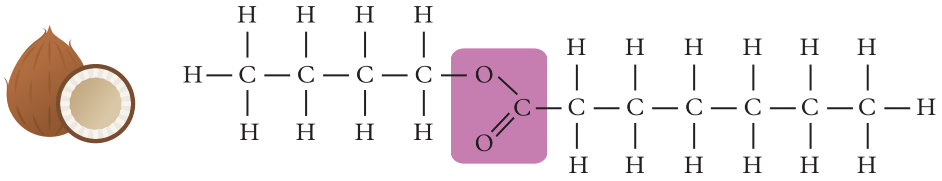


FIGURE 32 The ester functional group is in the middle of two alkyl groups (each is represented by the letter R).

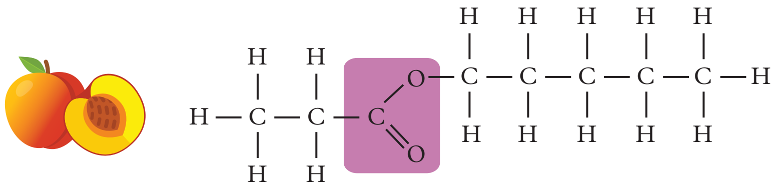
ester
a compound produced as the result of a reaction between an alcohol and a carboxylic acid



Methyl butanoate



Butyl heptanoate



Pentyl propanoate

FIGURE 33 Three esters and the scents they are responsible for. The ester functional groups are highlighted.

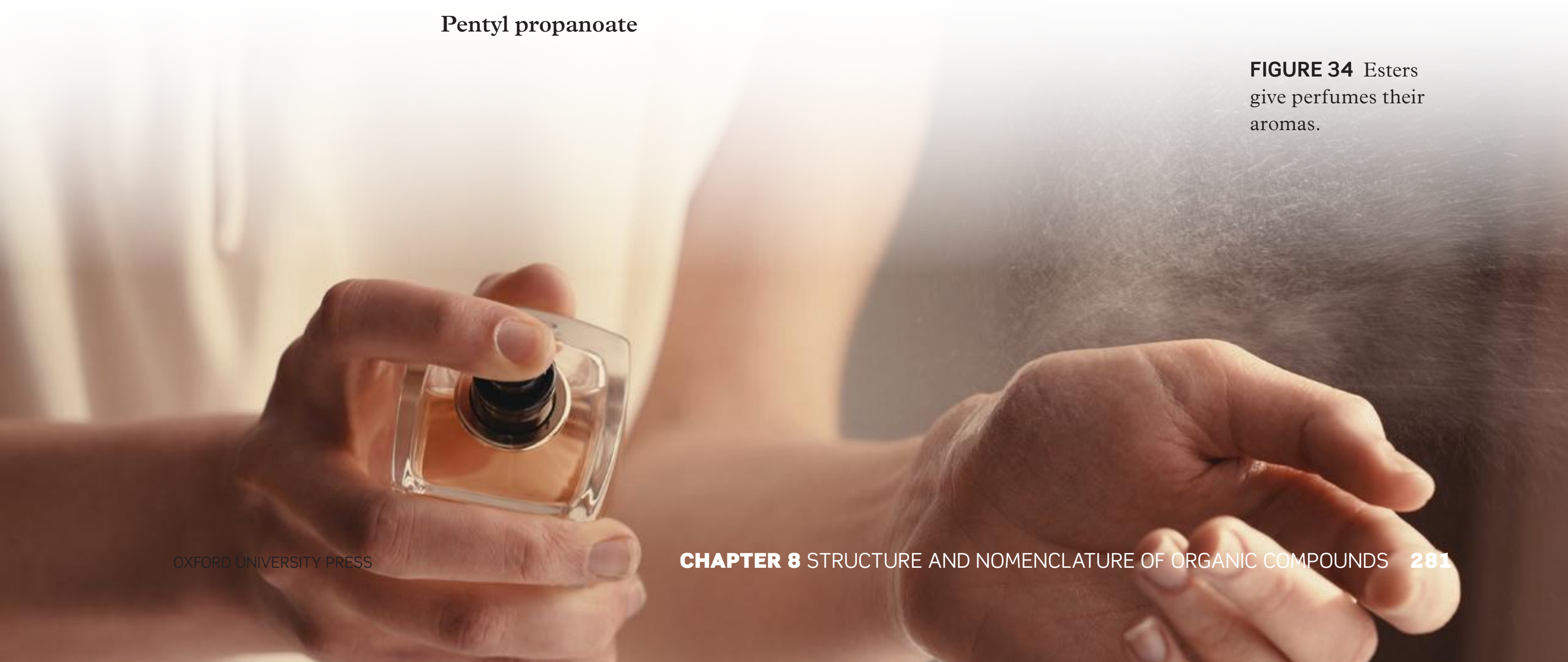


FIGURE 34 Esters give perfumes their aromas.

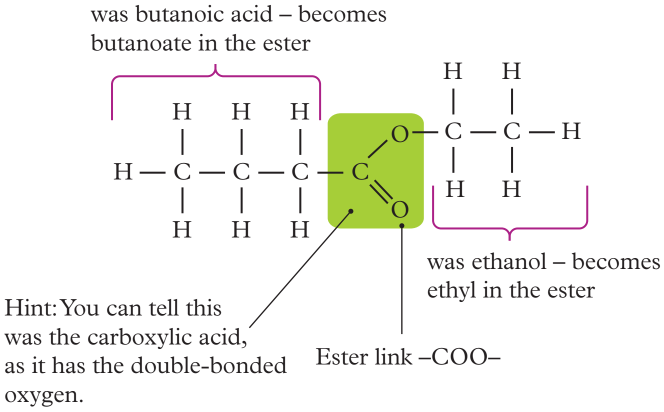
Naming esters

Naming esters is a bit different to naming other hydrocarbons with functional groups. The names are made up of the two original molecules that made it: an alcohol and carboxylic acid.

You can name esters using the following steps.

1. Name the alcohol first (the chain attached to the single-bonded oxygen). Keep the prefix and replace the '-anol' with an '-yl' suffix.
2. Name the carboxylic acid (the chain containing the carbonyl group). Keep the prefix and replace the '-anoic acid' with an '-oate' suffix.

For example, consider ethyl butanoate in Figure 35. The alcohol, attached to the single-bonded oxygen, has two carbon atoms, so it is ethyl. The carboxyl acid has four carbon atoms, including the carbon double-bonded to the oxygen, so it is butanoate. When we put these two together, we get ethyl butanoate.



was butanoic acid – becomes butanoate in the ester

was ethanol – becomes ethyl in the ester

Hint: You can tell this was the carboxylic acid, as it has the double-bonded oxygen.

Ester link –COO–

FIGURE 35 Ethyl butanoate

8.2 REAL-WORLD CHEMISTRY

Space smells

Space might seem like a whole lot of nothing, but it's actually full of stuff that smells!

Sagittarius B2 is the colossal molecular cloud of gas and dust located 390 light years from the centre of the Milky Way. In 2019, it hit the headlines because, according to scientists, it smells like raspberries.

What these scientists found inside Sagittarius B2 was huge quantities of a common ester you could make in a lab – ethyl methanoate (or ethyl formate, as it is commonly known). This ester is the chemical responsible for giving raspberries their flavour.

FIGURE 36 The cloud of dust and gas called Sagittarius B2 in the centre of our universe

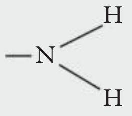
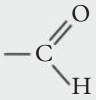
Apply your understanding

1. Identify the alcohol and carboxylic acid that make up the ester ethyl methanoate.
2. Draw the structure of ethyl methanoate.
3. Research why some chemicals have common names and why their IUPAC names are not always used.

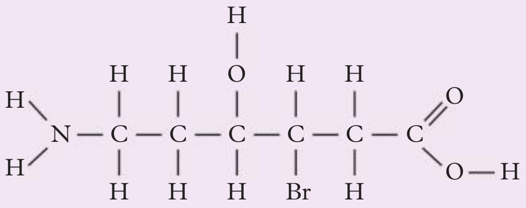
8.2 CHECK YOUR LEARNING

Describe and explain

1 Fill in the following table of functional groups.

Name	Functional group
Alkane	
	
Primary amide	
	—O—H
Ketone	
Carboxylic acid	
	C=C
	
Ester	
	–F, –Cl, –Br, –I

2 An organic compound is shown below.



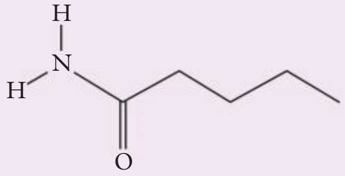
a Identify all the functional groups present in the compound.

b Identify the name of each functional group present.

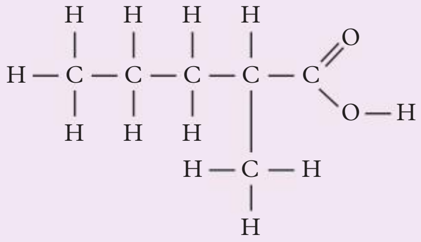
3 Identify the type of representation being used for the following compounds.

a $CH_3CH_2CH_2CH{=}CHCH_3$

b



c



4 Explain why benzene is not part of the alkene homologous series.

Apply, analyse and compare

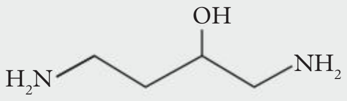
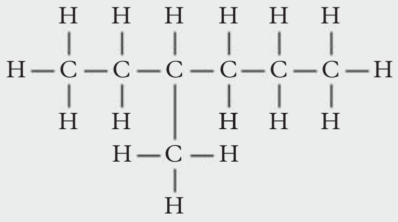
5 Compare cyclohexane to other alkanes.

6 Draw an amine and amide and analyse the structural differences in their functional groups.

7 Draw a primary, secondary and tertiary alcohol, and compare the differences in their structures.

8 Draw a ketone and an aldehyde and compare the differences in their structures.

9 Fill in the table below to demonstrate that you can convert between different molecular representations.

Semi-structural formula	Structural formula	Skeletal structure
$CH_3CH_2CHC(CH_3)CH_2OH$		
		
		

Design and discuss

10 The general formula for alkanes is C_nH_{2n+2}; for alkenes it is C_nH_{2n}. Consider the other homologous series we have covered, and design/create general formulas for each of them.

8.3

IUPAC systematic naming

KEY IDEAS

In this topic, you will learn that:

- naming hydrocarbons follows a set of rules stated by IUPAC
- functional groups have a priority order that affects the way organic molecules are named.

There is a set of rules for naming organic compounds in chemistry that are outlined by the International Union of Pure and Applied Chemistry (IUPAC). These rules exist so that chemists can consistently name, identify and communicate about organic structures.

Study tip

Between numbers and words, use a dash (e.g. 2-chloro). Between two numbers, use a comma (e.g. 2,2-dichloro). Don't put anything between words (e.g. 2-chloropentane).

IUPAC rules for naming organic compounds

You can name an organic molecule by following the steps below.

1. Determine the longest chain of carbon atoms – this will be your 'main chain'. Use the correct prefix based on the total number of carbon atoms (Table 1).
2. Determine which end is nearest to an alkyl branch, functional group, double or triple bond – this is your 'starting point'. The starting point will be the carbon atom closest to the highest priority group. You will learn about this on page 286.
3. Number the carbon atoms from the 'starting point' to the terminal carbon.
4. Identify and name any **alkyl group** branches (Table 2).
5. Identify any functional groups present and determine whether to use prefixes or suffixes depending on their type and position (Table 3).
6. When two or more branches or functional groups occur on the same or different carbon atom, the number of the carbon atom is indicated for each branch. Alkyl branches and halogens have the same priority and are named in alphabetical order.
7. When two or more identical branches or functional groups occur on different carbon atoms, the **counting prefixes** 'di-', 'tri-' and 'tetra-' are used.

alkyl group
the general name for a hydrocarbon branch consisting of any number of carbon atoms

counting prefixes
used to indicate multiple identical branches or functional groups; 2 = di-, 3 = tri-, 4 = tetra-

Naming the main chain

You should recall from Units 1 and 2 that the name of the main chain is based on the number of carbon atoms it contains. You can use Table 1 to help you remember the prefixes for each number of carbon atoms.

TABLE 1 Main carbon chain prefix

Number of carbon atoms	Prefix
1	Meth-
2	Eth-
3	Prop-
4	But-
5	Pent-
6	Hex-
7	Hept-
8	Oct-

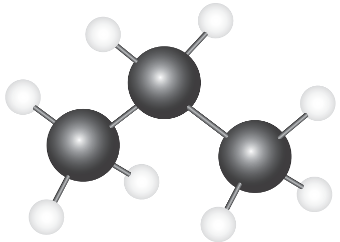


FIGURE 1 How many carbon atoms are in this compound?

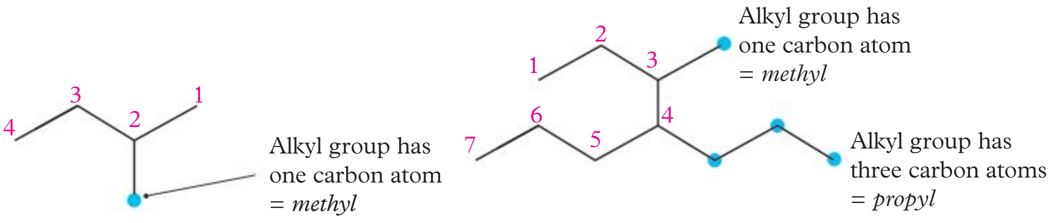
Naming branches

Molecules often have ‘branches’ off the main chain, called alkyl side chains or alkyl groups. Each alkyl group is given a name based on the number of carbon atoms in that branch (Table 2).

The branch name is placed before the molecule’s parent name and given a number to indicate where it is located along the main chain. If there is more than one branch, they are placed in alphabetical order (e.g. 4-ethyl-3-methylheptane).

TABLE 2 Branches are named based on their number of carbon atoms.

Number of carbon atoms in branch	Alkyl group
1	Methyl
2	Ethyl
3	Propyl
4	Butyl



2-methylbutane

3-methyl-4-propylheptane

FIGURE 2 **a** 2-methylbutane and **b** 3-methyl-4-propylheptane

Worked example 8.3A walks you through how to name branched hydrocarbons.

8.3A Worked example
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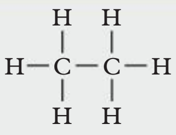
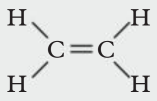
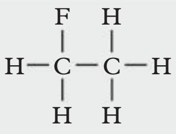
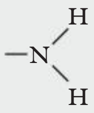
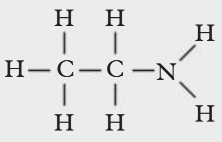
8.3A Worked example
Video demonstration

Naming the functional groups

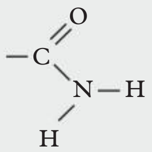
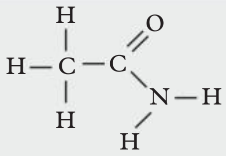
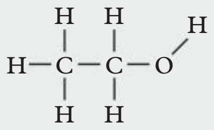
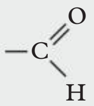
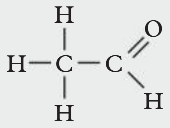
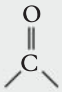
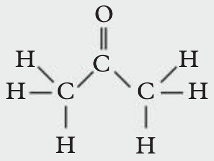
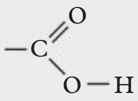
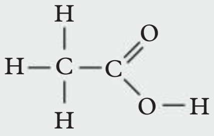
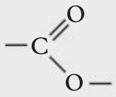
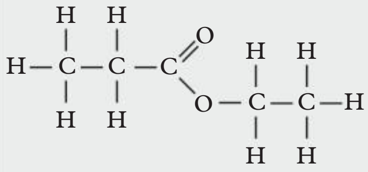
In Topic 8.2, you learnt about different types of functional groups and how to name them. This is summarised in Table 3.

TABLE 3 Functional group names and structures

Functional group	Prefix/ suffix	Homologous series	Example molecule
Carbon–carbon single bonds only C—C	-ane	Alkane	 Ethane
Carbon–carbon double bond C=C	-ene	Alkene	 Ethene
Halogen (group 17) atom (–F, –Cl, –Br, –I), collectively known as –X	fluoro-, chloro-, bromo-, iodo-	Haloalkane	 Fluoroethane
Amine group (–NH_2) on end 	-anamine amino-	Amine	 Ethanamine

(continued)

TABLE 3 continued

Functional group	Prefix/ suffix	Homologous series	Example molecule
Amide group ($-CONH_2$) on the end 	-anamide	Amide	 Ethanamide
Hydroxyl group $-OH$	-anol hydroxy-	Alcohol	 Ethanol
Carbonyl group on the end 	-anal formyl-	Aldehyde	 Ethanal
Carbonyl group in the middle 	-anone oxo-	Ketone	 Propanone
Carboxyl group 	-anoic acid	Carboxylic acid	 Ethanoic acid
Two carbon chains with an ester group ($-COO-$) in the middle 	-yl -anoate	Ester	 Ethyl propanoate

8.3B Worked example
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8.3B Worked example
Video demonstration

8.3C Worked example
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8.3C Worked example
Video demonstration

Worked examples 8.3B and 8.3C show you how to name organic compounds with multiple functional groups.

Priority naming

Sometimes organic molecules have multiple functional groups. When this happens, you need to work out which group to use as the ending of the name.

To do this, you need to consider the functional group priorities shown in the table below. The functional group with the highest priority will be the ending or suffix, and the other functional group(s) will all be prefixes to the name.

Like alkyl branches, halogens are named in alphabetical order. Halogens and alkyl branches have the same level of priority. Therefore, when a compound has both alkyl branches and halogens, they are all ordered alphabetically.

Study tip

When naming organic compounds with multiple functional groups as prefixes, make sure they are listed in alphabetical order. Counting prefixes are not included in the alphabetical listing. For example, in 2-bromo-3,4-dimethylhexane, 'di' doesn't change the order. *Bromo* comes before *methyl* alphabetically.

TABLE 4 Functional group properties

Priority	Functional group	Suffix	Alternative name/prefix
Highest	Carboxylic acid –COOH	-oic acid	
	Amide $–CONH_2$	-amide	
	Aldehyde –CHO	-al	formyl-
	Ketone –CO	-one	oxo-
	Alcohol –OH	-ol	hydroxy-
	Amine $–NH_2$	-amine	amino-
	Alkene C=C	-ene	-en-
	Alkane C–C	-ane	-an-
Lowest	Halo –F, –Cl … etc.		fluoro- chloro-

Bringing it all together

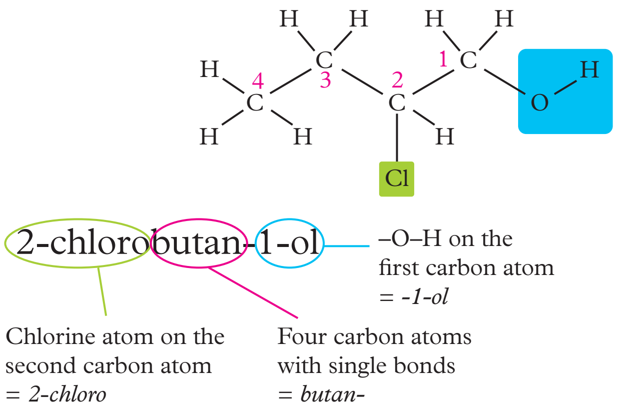
Let's look at an example to put these rules into practice. The following can be observed from Figure 3:

- Different parts of the name correspond to different parts of the compound.
- The main carbon chain of the molecule is given a name based on the number of carbon atoms it contains. Here we have four carbon atoms, so the prefix is 'but-' (see Table 1).
- There are two functional groups in the molecule in Figure 3: a chlorine on the second carbon atom, and a hydroxyl group on the first. We therefore write 2 before 'chloro-' and 1 before '-ol' to identify where each functional group is located on the main chain.

The worked examples in this topic will show you how to apply these rules to a diverse range of molecules.

Study tip

When naming organic compounds, if a number is not required and you add it into the name, it will be marked as incorrect. E.g., ethan-1-ol is just ethanol, as the –OH has to be on a 1–C, because there are no other options for it to be on. The only correct answer is therefore *ethanol*.



–O–H on the first carbon atom = *-1-ol*

Chlorine atom on the second carbon atom = *2-chloro*

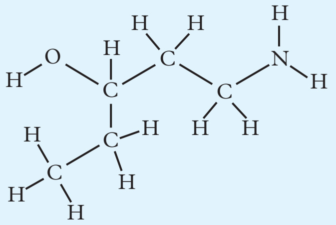
Four carbon atoms with single bonds = *butan-*

FIGURE 3 Naming of organic compounds using IUPAC rules: 2-chlorobutan-1-ol is colour coded so that you can see what each part of the name relates to.

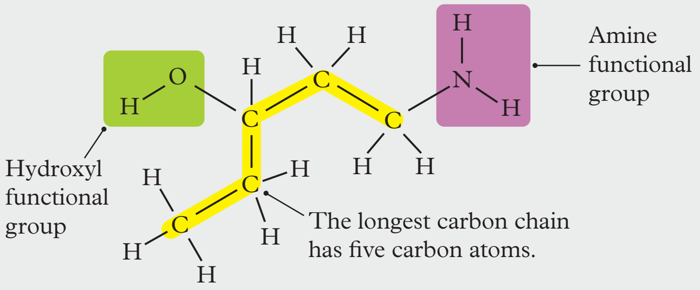
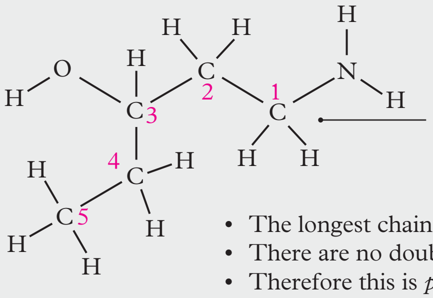
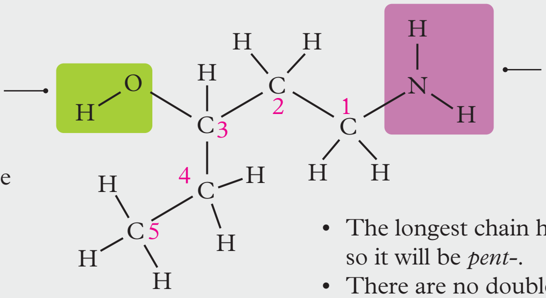
8.3D WORKED EXAMPLE

USING PRIORITY NAMING RULES TO NAME AN ORGANIC COMPOUND WITH MULTIPLE FUNCTIONAL GROUPS

Give the systematic name of the compound shown.



Solution

Think	Do
Step 1: Identify the longest chain and any double bonds and functional groups. Also look for the end with the functional group closest to it – this will be the 'starting point'. Start numbering from that end.	 Amine functional group Hydroxyl functional group The longest carbon chain has five carbon atoms. Note: The –OH functional group is of higher priority than the –NH functional group, so it is an alcohol with an amino functional group.
Step 2: Number the carbon atoms from the starting point. In this case, the highest-priority functional group is on the third carbon atom (in the middle), so start numbering from the end with the next highest functional group.	 Start numbering from this carbon atom. • The longest chain has five carbon atoms so it will be *pent-*. • There are no double bonds in the molecule, so it will be *-an-*. • Therefore this is *pentan-*.
Step 3: Now that the carbon atoms have been numbered, identify the functional group and double bond and determine names and numbers.	The hydroxyl group is on the third carbon atom but has the highest naming priority, so it is the suffix *-3-ol*.  The amine is on the first carbon, but is lower in naming priority to the alcohol, so it's the prefix *1-amino*. • The longest chain has five carbon atoms so it will be *pent-*. • There are no double bonds in the molecule, so it will be *-an-*. • Therefore this is *pentan-*.
Step 4: Combine everything together in the correct order.	The systematic name of this compound is 1-aminopentan-3-ol.

Up for a challenge? Have a go at naming and drawing all the isomers for pentanoic acid in Challenge 8.3.

8.3 Challenge
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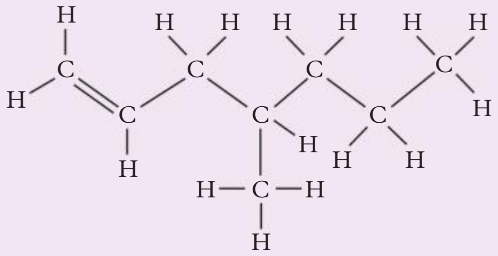
8.3 CHECK YOUR LEARNING

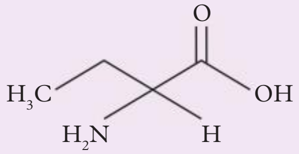
Describe and explain

1 Four organic compounds I–IV are shown below.
  - **a** Identify the type of representation used for each of the molecular structures shown.
  - **b** Name the molecules.

I



II



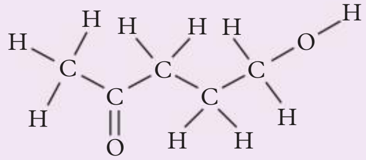
III $CH_3CH_2CH_2COOCH_2CH_3$

IV

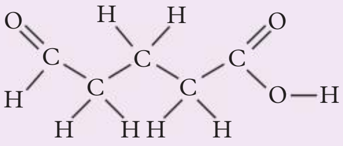
2 Draw the structural formulas from the names of the following organic molecules.
  - **a** 2-methylpropanoic acid
  - **b** 3-ethylpent-2-ene
  - **c** Ethan-1,2-diol
  - **d** Butyl propanoate
  - **e** Pent-4-en-1-ol

3 Highlight and name the functional groups in the molecules shown. For each molecule, identify the highest priority functional group and the suffix of the molecule.

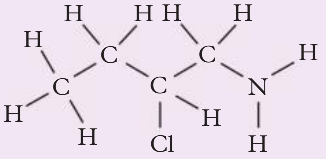
a



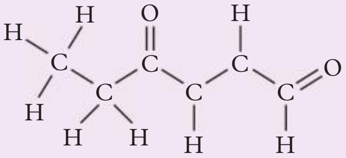
b



c



d



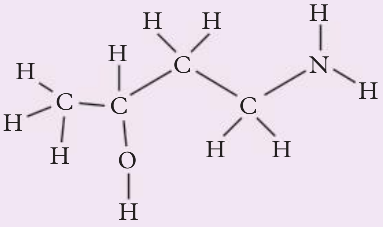
Apply, analyse and compare

4 Apply the IUPAC naming rules to the molecules from Question 3.

5 Name and draw the ester made from an alcohol with six carbon atoms and a carboxylic acid with two carbon atoms.

Design and discuss

6 Discuss why the molecule shown is called 4-aminobutan-2-ol and not 1-aminobutan-3-ol.



7 Discuss why it is important to have a priority of functional groups when naming hydrocarbons.

8.4 Physical properties of organic compounds

KEY IDEAS

In this topic, you will learn that:

- physical properties of organic molecules include boiling point, melting point and viscosity
- functional groups impact the physical properties of organic compounds
- chain length and branching will also impact the physical properties of organic compounds.

In this topic, we will examine the physical properties of organic molecules including:

- boiling point
- melting point
- viscosity.

The functional groups that we have learnt about in this chapter play an important role in the physical properties of organic compounds. This is due to **intermolecular forces** between the molecules. From Units 1 and 2, recall the intermolecular forces that are present between organic molecules:

intermolecular forces
the attractive forces between covalently bonded molecules, including hydrogen bonding, dipole–dipole attractions and dispersion forces

Strongest ↓ Weakest

- Hydrogen bonding
- Dipole–dipole attraction
- Dispersion forces

Hydrogen bonding is the strongest of the intermolecular forces, followed by dipole–dipole attraction and finally dispersion forces. The strength of these intermolecular forces is important, as it is one of the key factors affecting the physical properties of molecules.

Study tip

Remember that for a molecule to form hydrogen bonds, it has to have a 'FON' and a hydrogen. FON can be any one of fluorine, oxygen and nitrogen, and that FON atom has to have unpaired valence electrons and a hydrogen attached to it.

Boiling points of organic compounds

The **boiling point** of a liquid is the temperature at which it changes from its liquid state into a gas state. For a compound to change its state, enough energy needs to be applied to disrupt the intermolecular forces between its molecules.

boiling point
the temperature at which a liquid becomes a gas

Key factors that contribute to the boiling point of organic compounds include:

- the carbon chain length or molecule size
- the molecular shape and amount of branching in the molecule
- the functional group(s) present.

These factors determine the intermolecular forces between molecules. Let's consider each in detail.

TABLE 1 Alkanes and their boiling points

Alkane	Boiling point (°C)
Methane	–161.6
Ethane	–89
Propane	–42
Butane	–1
Pentane	36.1
Hexane	69

The effect of carbon chain length or molecule size on the boiling point

The length of the carbon chain, or molecule size, impacts the boiling point of organic molecules. Table 1 outlines the first six alkanes and their boiling points.

As the chain length increases, the size of the molecule increases, and therefore the amount of dispersion forces between chains also increases.

8.4
PRACTICAL: MODELLING

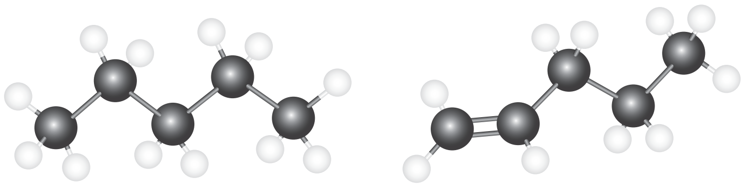
How do the shapes of molecules affect their intermolecular bonding and physical properties? Go to your **Student obook pro**.

This increase in the number of dispersion forces means that more energy will be required to break or disrupt the dispersion forces between molecules. If more energy is required to break intermolecular forces, then the molecule will have a higher boiling point.

The effect of the molecular shape and amount of branching on the boiling point

The molecular shape of an organic compound, such as the presence of multiple bonds and branches, will also impact the boiling point.

Alkenes have a lower boiling point than alkanes. This is because of the double bond, which stops atoms getting as close to each other, meaning the dispersion forces won't be as strong in alkenes with the same number of carbon atoms. Alkanes, which are straight chain hydrocarbons, can pack in together very easily.



Pentane
Boiling point = 36.1°C

Pent-1-ene
Boiling point = 30°C

FIGURE 1 The boiling points and structures of pentane and pent-1-ene

In Figure 1, the double bond in pent-1-ene changes the shape of the molecule so that its atoms cannot pack in as tightly as the pentane molecule's atoms. As a result, the boiling point of pent-1-ene is lower.

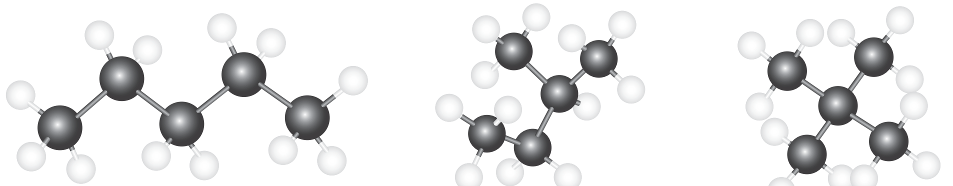
The effect of branches on the boiling point

The boiling point of molecules is also impacted by branches. As the number of branches increase, the boiling point decreases. This is because there is less surface area and therefore fewer places for the weak dispersion forces to form between the molecules.

The molecules are also further apart from each other because the branches get in the way of the molecules being able to pack together. This also decreases the boiling points of branched molecules (Figure 2).

Study tip

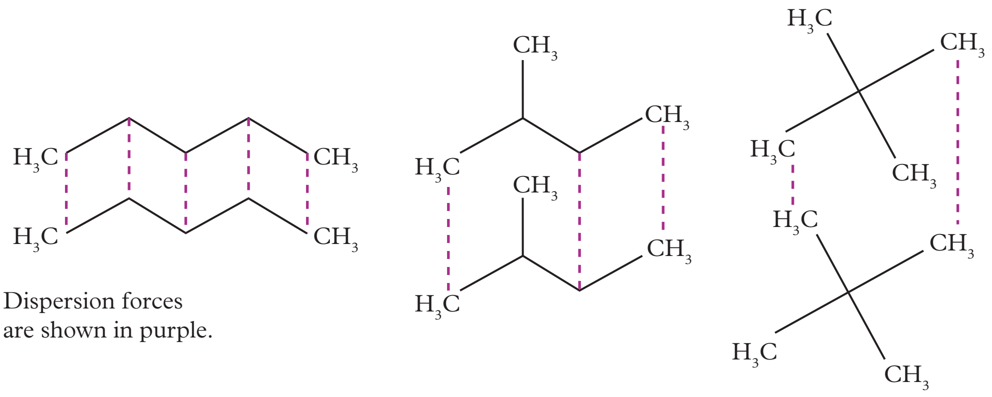
The best way to see molecular shape and how it affects the packing of molecules is to build the molecules. If your school doesn't have molecular model kits, you can use household items instead.



Pentane
Boiling point = 36.1°C

2-methylbutane
Boiling point = 27.8°C

2,2-dimethylpropane
Boiling point = 9.5°C



Dispersion forces are shown in purple.

FIGURE 2 How branching affects the boiling point of the branched isomers of pentane

The effect of functional groups on the boiling point

Functional groups also influence the boiling point of organic compounds. This is because the intermolecular forces of each functional group vary, and therefore the amount of energy required to break their bonds varies. The stronger the intermolecular forces, the more energy is required to disrupt them, and therefore the higher the boiling point.

The intermolecular forces present in all the functional groups we explored in Topic 8.2 are shown in Table 2.

TABLE 2 Functional groups and their intermolecular forces

Functional group	Representation	Strongest intermolecular force present between molecules
Alkane	C–C	Dispersion forces
Alkene	C=C	Dispersion forces
Halogen	–F, –Cl, –Br, –I	Dipole–dipole attraction
Amine	$-NH_2$	Hydrogen bonding
Amide	$-CONH_2$	Hydrogen bonding
Alcohol	–OH	Hydrogen bonding
Ketone	–CO	Dipole–dipole attraction
Aldehyde	–CHO	Dipole–dipole attraction
Carboxylic acid	–COOH	Hydrogen bonding
Ester	–COO–	Dipole–dipole attraction

The molecules with hydrogen bonding or a dipole–dipole attraction will also have dispersion forces between the hydrocarbon chains. We discuss hydrogen bonds and dipole–dipole attractions more often than dispersion forces, because they are stronger and have a greater effect on the attraction between the molecules.

Let’s compare propane and propanol to demonstrate this concept. Look at Figure 3. Propane only has dispersion forces between molecules. Propanol has an –OH functional group and therefore has hydrogen bonding as well as dispersion forces between molecules. Hydrogen bonding is a stronger intermolecular force, so more energy is required to disrupt the bonds between propanol molecules. Propanol therefore has a higher boiling point than propane (97°C compared to –42°C).

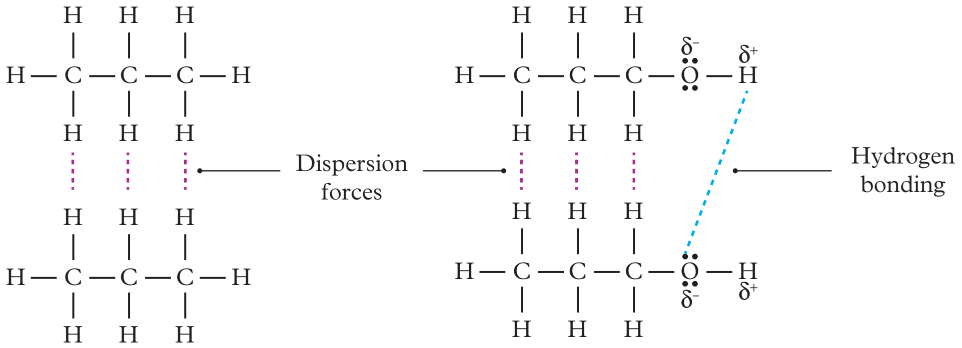
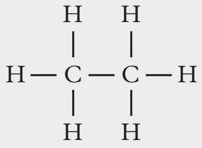
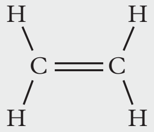
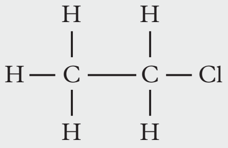
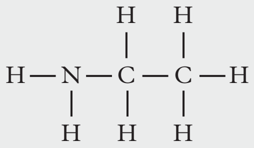
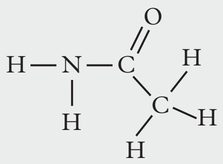
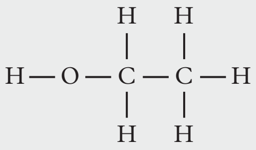
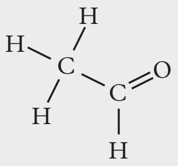
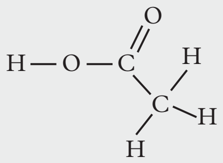
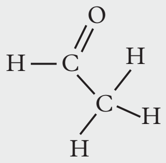


FIGURE 3 The intermolecular forces between propane and propanol are dispersion forces and hydrogen bonding.

This makes propane gas, and propanol a liquid at SLC (25°C and 100 kPa).

Table 3 shows a comparison of some molecules of a similar size. Here you can see the relationship between functional groups, intermolecular bonding and boiling point.

TABLE 3 The boiling points and intermolecular forces of two carbon chain molecules with different functional groups. Note: there is no ketone with two carbon atoms, so no ketone is included in this table.

Name	Structure	Strongest intermolecular force present	Boiling point
Ethane	 $\mathrm{H{-}\underset{H}{\overset{H}{C}}{-}\underset{H}{\overset{H}{C}}{-}H}$	Dispersion forces	–89°C
Ethene		Dispersion forces	–103.7°C
Chloroethane	 $\mathrm{H{-}\underset{H}{\overset{H}{C}}{-}\underset{H}{\overset{H}{C}}{-}Cl}$	Dipole–dipole attraction	12.3°C
Ethanamine	 $\mathrm{H{-}\underset{H}{N}{-}\underset{H}{\overset{H}{C}}{-}\underset{H}{\overset{H}{C}}{-}H}$	Hydrogen bonding	16.6°C
Ethanamide		Hydrogen bonding	222°C
Ethanol	 $\mathrm{H{-}O{-}\underset{H}{\overset{H}{C}}{-}\underset{H}{\overset{H}{C}}{-}H}$	Hydrogen bonding	78.37°C
Ethanal		Dipole–dipole attraction	20.2°C
Ethanoic acid		Hydrogen bonding	118°C
Methyl methanoate		Dipole–dipole attraction	31.8°C

Boiling points of haloalkanes

Haloalkanes have higher boiling points than alkanes because of their stronger dipole–dipole forces.

In Table 3, you will notice that chloroethane has a significantly higher boiling point than ethane (12.3°C compared to –89°C). This is due to the polar carbon–chlorine bond in chloroethane, which forms a permanent dipole (Figure 4). This then allows the formation of dipole–dipole attractions between chloroethane molecules.

Dipole–dipole attractions are much stronger intermolecular forces than dispersion forces. Because the intermolecular force is much stronger, it will take more energy for those bonds to be disrupted, which makes the boiling point of the haloalkanes higher than that of alkanes with the same number of carbon atoms.

The trend of boiling points within the haloalkanes is relative to atomic size, not electronegativity. Looking at Table 4, you can see that the boiling point of the haloalkanes increases as the size of the halogen attached to the hydrocarbon increases.

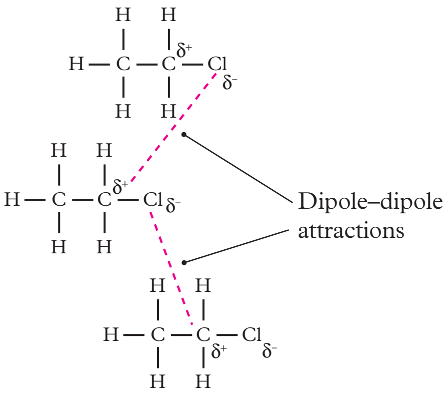
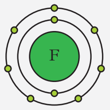
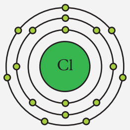
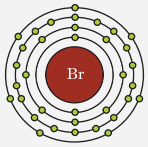
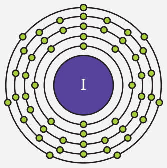


FIGURE 4 The carbon–chlorine bond and dipole–dipole interactions between chloroethane molecules

TABLE 4 Haloalkanes and their boiling points

Electronegativity decreases, but atomic size increases.

Haloalkane	Halogen size	Boiling point
Fluoroethane		–37.1°C
Chloroethane		12.3°C
Bromoethane		38°C
Iodoethane		72°C

Boiling points of amines, alcohols, carboxylic acids and amides

Alcohols, amines, carboxylic acids and amides all have the ability to form hydrogen bonds. Hydrogen bonding is the strongest of all the intermolecular forces, so it makes sense that these molecules have higher boiling points than those with only dispersion forces or dipole–dipole attractions.

From Table 3, you can see the trend in boiling points is: ethanamine 16.6°C < ethanol 78.37°C < ethanoic acid 118°C < ethanamide 222°C.

All four have hydrogen bonding between their molecules, but there is a large difference between their boiling points – let's understand why.

Primary amines and alcohols

Amines and alcohols both have polar functional groups so will both form hydrogen bonds, as shown in Figure 5 and Figure 6. The strength of attraction between the hydroxyl groups in alcohols is stronger than that between the amine groups because nitrogen is less electronegative than oxygen. This means that alcohols will have higher boiling points than amines with the same number of carbon atoms.

Primary alcohols have a higher boiling point than secondary and tertiary alcohols because of the placement of their hydroxyl functional group (Figure 7).

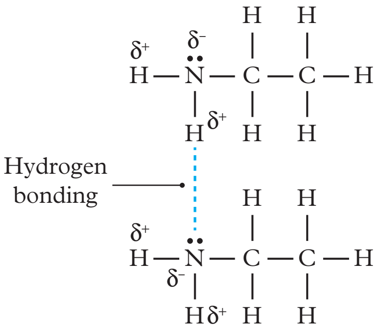


FIGURE 5 Hydrogen bonding between two ethanamine molecules due to the polar amide group

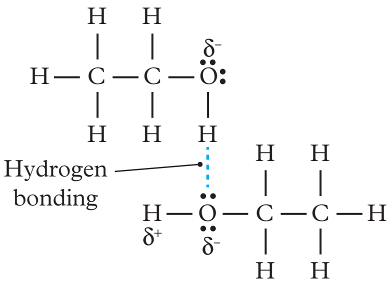
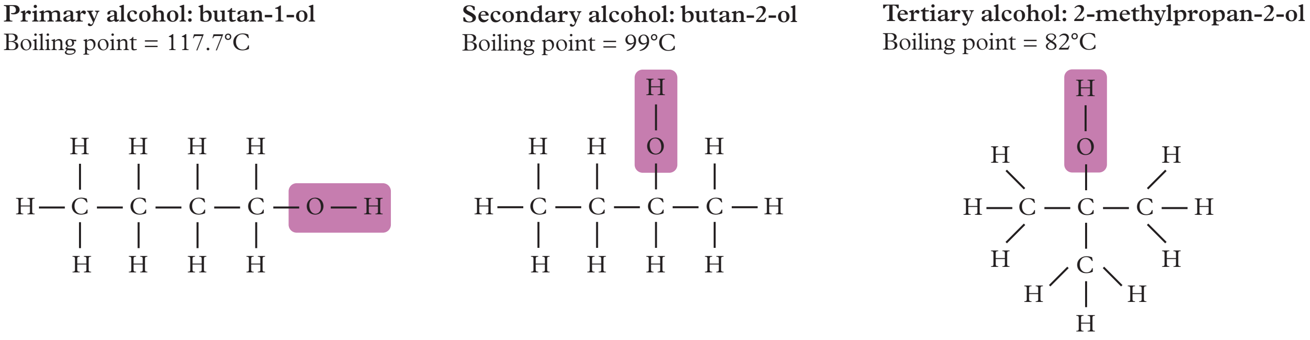


FIGURE 6 Hydrogen bonding between two ethanol molecules due to polar functional groups



Primary alcohol: butan-1-ol
Boiling point = 117.7°C

Secondary alcohol: butan-2-ol
Boiling point = 99°C

Tertiary alcohol: 2-methylpropan-2-ol
Boiling point = 82°C

FIGURE 7 Structures and boiling points of the primary, secondary and tertiary alcohols of C_4H_9OH

Propan-1-ol has a higher boiling point than the secondary alcohol propan-2-ol and the tertiary alcohol 2-methylpropan-2-ol. This is because the hydroxyl groups in the secondary and tertiary alcohols are more crowded than the hydroxyl group on the terminal carbon atom of a hydrocarbon chain. The crowding of the hydroxyl group makes it more difficult for hydrogen bonds to form.

The hydrogen bonds are not as strong and less energy is required to break the molecules for secondary and tertiary alcohols. They therefore have lower boiling points than primary alcohols.

The same goes for the placement of amines. If you consider propan-1-amine, its boiling point of 47.8°C is higher than that of its isomer propan-2-amine, which is 34°C. Again, this is because of the crowding of the functional groups, which restricts intermolecular bonding.

Carboxylic acids

Carboxylic acids also have hydrogen bonding between molecules, but, as you can see from Table 3 (ethanoic acid), they have a much higher boiling point than alcohols of the same size. This is because they form **dimers**.

dimer
two monomers joined by intermolecular bonds

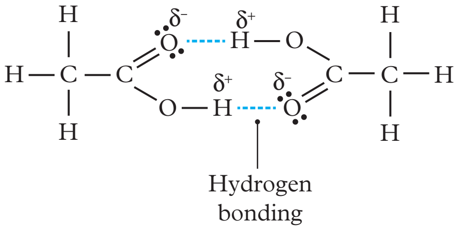


FIGURE 8 Two carboxylic acid molecules forming a dimer

A dimer between carboxylic acids forms due to hydrogen bonding between a double-bonded oxygen in one carboxyl group and the hydrogen from the other carboxyl group, and vice versa.

The dimer in Figure 8 is considered a larger, single molecule with a higher molar mass. Because the molecular size is increased, the dispersion forces between dimers of carboxylic acid pairs increase. More energy is required to break the intermolecular forces, so the boiling point is higher.

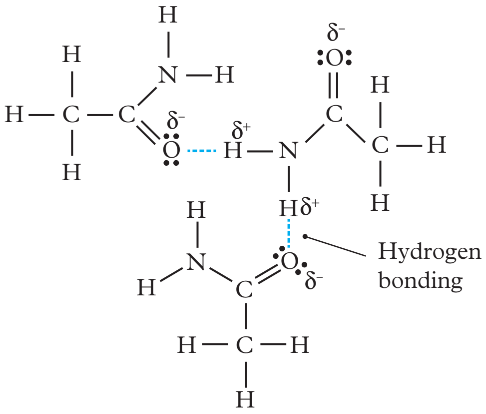


FIGURE 9 A primary amide can form multiple hydrogen bonds with other primary amides.

Primary amides

Primary amides have the highest boiling points of all the functional groups. They can form mutiple hydrogen bonds with other primary amide molecules (Figure 9).

The more hydrogen bonds between the molecules, the more energy it will take to break these bonds and therefore the higher the boiling point.

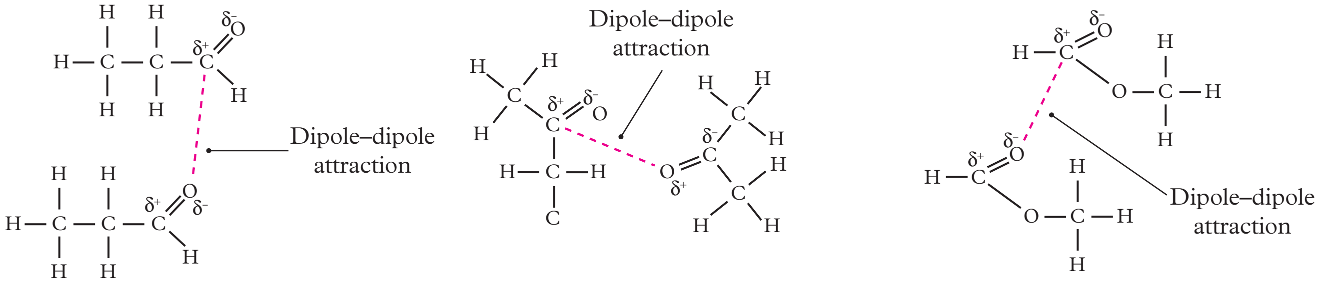
Boiling points of aldehydes, ketones and esters

Aldehydes, ketones and esters all have a double-bonded oxygen atom in their structure. Due to the difference in electronegativity between the double-bonded oxygen atom and the carbon atom it is attached to, aldehydes, ketones and esters have a permanent dipole that can attract nearby molecules.

The dipole–dipole attractions influence boiling points of molecules containing these functional groups. Their boiling points are lower than those of alcohols of the same size, but higher than those of alkenes of the same size.

Esters, ketones and aldehydes of the same size all have similar boiling points (Figure 10). There are some differences between them due to:

- the position of the carbonyl group
- the difference in electronegativity between the oxygen atom and the carbon atom it is attached to
- the shape of the molecule.



Propanal
Molar mass = 58 g mol^{-1}
Boiling point = 49°C

Propanone
Molar mass = 58 g mol^{-1}
Boiling point = 56°C

Methyl methanoate
Molar mass = 60 g mol^{-1}
Boiling point = 31.8°C

FIGURE 10 The permanent dipoles formed and dipole–dipole bonding between two molecules of similar-sized aldehydes, ketones and esters, and their boiling points

Melting points of organic compounds

Melting point is the temperature at which a substance changes from a solid to a liquid. The trend for melting points in organic molecules is the same as that for boiling points, but the melting points for substances are a lot lower than boiling points.

melting point
the temperature at which a solid becomes a liquid

All of the first eight alkanes have melting points under 0°C (Figure 11). This means that they are solids at room temperature. It is not until alkanes have above 16 carbon atoms in length that they melt and form liquids around room temperature.

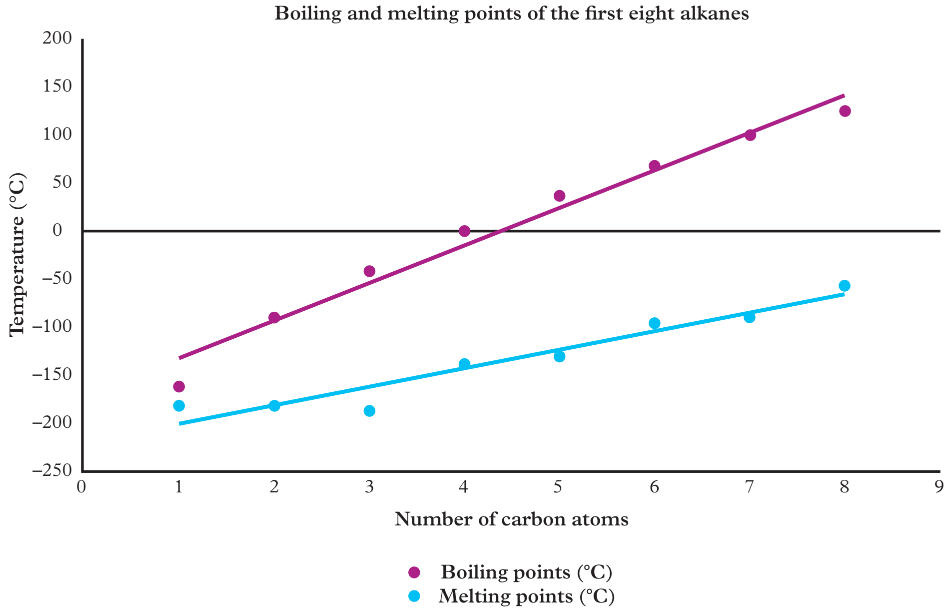


FIGURE 11 Boiling points and melting points of the first eight alkanes

Viscosity of organic compounds

Viscosity is the measure of the resistance of flow of a substance. A substance with high viscosity will flow slower than a substance with low viscosity. Viscosity is dependent on the intermolecular forces between the molecules in the substance (Figure 12).

viscosity
the ability of a substance to flow

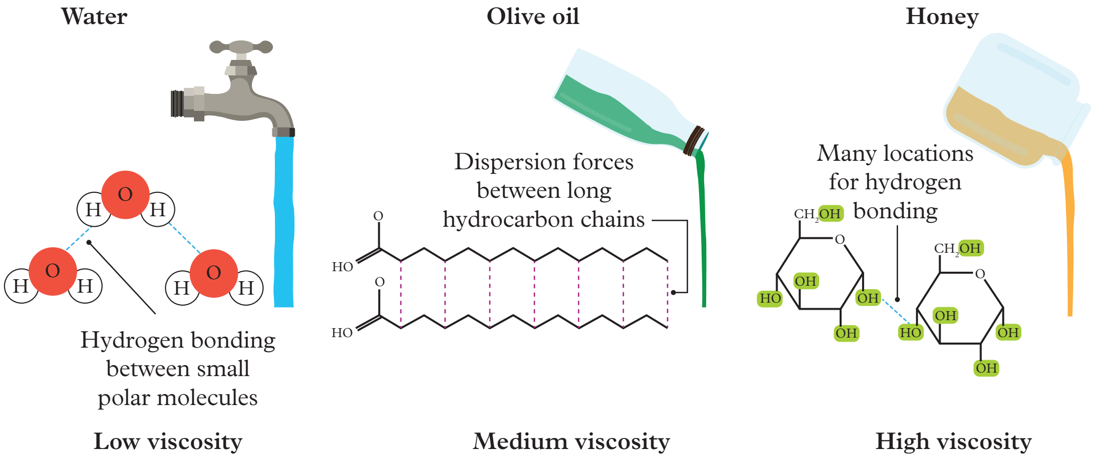


FIGURE 12 The intermolecular bonding between molecules within water, olive oil and honey

Generally, the stronger the bonding between molecules, the higher the viscosity of the substance. Honey and water both have hydrogen bonding between the molecules, but water is made up of small molecules of hydrogen and oxygen. Honey has a higher viscosity than water, as it is a larger molecule and has many hydroxyl groups that form hydrogen bonds with other hydroxyl groups. This means that the viscosity of honey is higher than that of water.

Olive oil contains long chain hydrocarbons that only have dispersion forces between the molecules. But as the chain length of hydrocarbons increases, so do the dispersion forces between molecules and therefore the viscosity. The inverse of this is also true: the smaller the molecule, the weaker the dispersion forces between molecules and therefore the lower the viscosity.

Temperature also has an effect on the viscosity of a substance. If you increase the temperature of a substance, you add energy and therefore weaken or break some of the intermolecular forces between the molecules. This decreases the viscosity of the substance, causing it to flow more freely. Cooling a substance will have the opposite effect on viscosity and make it harder for the substance to flow.

8.4 REAL-WORLD CHEMISTRY

The changing viscosity of frog saliva

Frog saliva is 175 times more viscous than human saliva. Frogs can change the viscosity of their saliva, which has helped them evolve to be insect-catching champions.

Frogs use their tongues to snatch insects from the air in under a second. In fact, the speed of their tongues is about $4\,\mathrm{m\,s^{-1}}$. This quick and powerful movement turns the frog's normally thick and viscous saliva into a thin water-like liquid that is quickly able to coat and trap the insect. The saliva then thickens again as the tongue reels back into the frog's mouth, together with its prey.

But how does the frog get the insect off its tongue, if it's stuck to its saliva? The frog retracts its big bulbous eyeballs into its skull and mouth cavity to push the insect off the tongue and into its mouth.

So the next time you want to catch a fly with your tongue, maybe try coating it in honey first … it may make the fly taste better too!

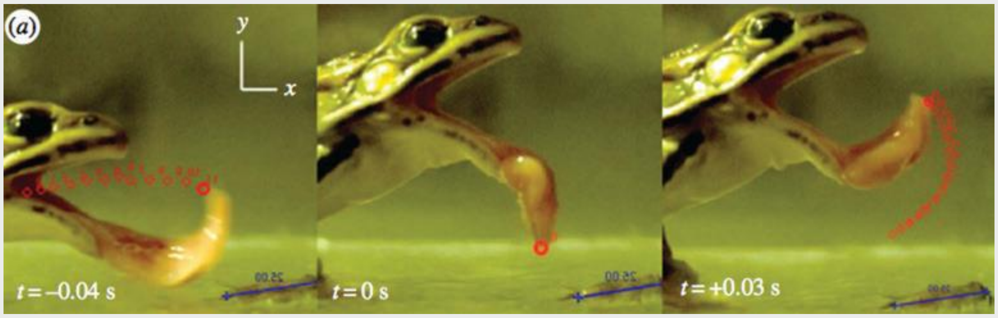


FIGURE 13 Frog snatching an insect out of the air in under a second

Apply your understanding

1 Define the term *viscosity*.

2 Describe how the change in viscosity of a frog's saliva helps it to catch insects.

3 A human tongue can't apply the same force as that of a frog, but the viscosity of human saliva could be changed in a similar way. Discuss how this can be done.

8.4 CHECK YOUR LEARNING

Describe and explain

1 Identify the types of intermolecular bonding in these molecules and explain the effect of the intermolecular bonding on their boiling points.
  - **a** Chloromethane
  - **b** Methanamine
  - **c** Methanal
  - **d** Methanol
  - **e** Methanoic acid

2 Describe the effect of intermolecular bonding on the boiling point.

3 Explain the differences between melting points and boiling points of organic compounds.

4 Describe why a tertiary alcohol will have a lower boiling point than a primary or secondary alcohol of the same size.

Apply, analyse and compare

5 Analyse the boiling points of the following molecules and put them in order of lowest to highest.
  - **a** Propene
  - **b** Pentane
  - **c** Octane
  - **d** Hexene
  - **e** 2-methylbutane

6 Analyse the boiling points of the following molecules and put them in order of lowest to highest.
  - **a** 1-chloropentane
  - **b** Pentan-2-one
  - **c** Pentan-1-ol
  - **d** Pentan-1-amine
  - **e** Pentanoic acid
  - **f** Pentanal
  - **g** Pentan-2-ol
  - **h** Pentan-3-one
  - **i** Pentanamide

7 Three molecules are listed below.
  - I Butan-1-ol
  - II Hex-3-ene
  - III Pentan-1,2,4-triol
  - **a** Draw the structural formula for each compound.
  - **b** Identify the functional groups and intermolecular forces between each.
  - **c** Compare the compounds and order them from:
    - **i** highest to lowest boiling point
    - **ii** highest to lowest viscosity.

Design and discuss

8 Butan-1-amine has a molar mass of 73 g mol^{-1} and a boiling point of 78°C; propanoic acid has a molar mass of 74 g mol^{-1} and a boiling point of 141.2°C.

Discuss why there is such a difference in boiling points between the two molecules even though they have similar molar mass.

CHAPTER 8 Review

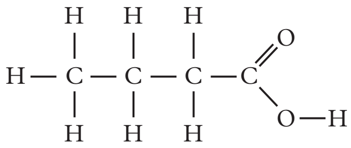
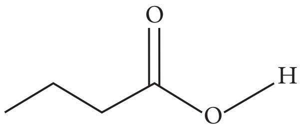
Chapter summary

8.1

- Carbon can form many diverse compounds because of its four valence electrons.
- Carbon can form single, double and triple bonds with itself and other atoms.
- Bond strength is dependent on bond energy and is measured in $kJ\,mol^{-1}$. The higher the bond energy, the more stable the bond and therefore the molecule.
- Bond length is the distance between the nuclei of the two atoms involved in the bond. As you move across a period in the periodic table, the bond length decreases. As you move down a group, the bond length increases. Distance between single bonds is greater than distance between double bonds involving the same molecules.
- Structural isomers can be positional, chain or both; they have the same molecular formulas but different structures.

8.2

- There are four ways of representing organic molecules.

Formula	Representation
Molecular formula	$C_4H_8O_2$
Structural formula	
Semi-structural (condensed) formula	$CH_3CH_2CH_2COOH$ or $CH_3(CH_2)_2COOH$
Skeletal structure	

- Alkanes are saturated hydrocarbons and contain only C–C single bonds.
- Alkenes are unsaturated hydrocarbons and contain at least one C=C double bond.
- Cyclohexane is a saturated cyclic hydrocarbon. It does not have any terminal carbon atoms.
- Benzene is an unsaturated cyclic hydrocarbon made up of six carbon atoms, with three carbon–carbon double bonds in a ring structure.
- Functional groups are an atom or group of atoms that give specific characteristics to molecules.
- Haloalkanes are hydrocarbons with a halogen (F, Cl, I, Br) functional group.
- Primary amines have a nitrogen atom with two hydrogen atoms bonded to a carbon atom along the main chain of an organic molecule.
- Primary amides occur where a terminal carbon atom has a double-bonded oxygen atom and a nitrogen atom with two hydrogen atoms attached to it.
- Primary alcohols are a functional group containing an oxygen atom and a hydrogen atom (hydroxyl group) attached to a terminal carbon atom. Secondary alcohols are located along the carbon chain.

8.2 Tertiary alcohols contain a hydroxyl group that is attached to a carbon which is connected to two more carbon atoms.

- Carboxylic acids have a terminal carbon atom with a double-bonded oxygen atom and a hydroxyl group attached.
- An ester is formed from a reaction between an alcohol and a carboxylic acid and has a –COO– along the chain.

8.3
- IUPAC naming rules are used to name organic molecules.
- If a compound has multiple functional groups, molecules should be named based on priority naming order.

8.4
- As chain length increases, intermolecular bonding between molecules increases, so the boiling point, melting point and viscosity increase.
- The more branches on a chain, the weaker the intermolecular bonding and therefore the lower the boiling point, melting point and viscosity.
- The stronger the intermolecular forces, the higher the boiling point, melting point and viscosity.
- Increasing the temperature of a substance will decrease its viscosity and allow it to flow more readily.

Key formulas

Degree of unsaturation	$\text{DoU} = \frac{2C + 2 + N - X - H}{2}$

Chapter checklist

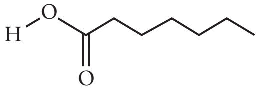
Use the success criteria in the table below to rate how well you understand each concept as ‘Confidently’, ‘Mostly’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Mostly	Not really	Revision links
• explain how the carbon atom contributes to the diversity of organic compounds, including: valence electron number, relative bond strength, relative stability of carbon bonds with other elements, degree of unsaturation, the formation of structural isomers	☐	☐	☐	Go back to Topic 8.1
• draw and convert between molecular, structural, semi-structural formulas and skeletal structures of organic compounds	☐	☐	☐	Go back to Topic 8.2
• describe the structure of the following homologous series and functional groups: alkanes (including cyclohexane), alkenes, benzene, haloalkanes, primary amines, primary amides, alcohols (primary, secondary and tertiary), aldehydes, ketones, carboxylic acids and non-branched esters	☐	☐	☐	Go back to Topic 8.2
• identify the IUPAC name of organic compounds with branches and up to two functional groups	☐	☐	☐	Go back to Topic 8.3
• draw structures of organic compounds based on the IUPAC name	☐	☐	☐	Go back to Topic 8.3
• explain how the structure and bonding of homologous series affect boiling point, melting point and viscosity	☐	☐	☐	Go back to Topic 8.4

Revision questions

Multiple choice

1 Identify the correct IUPAC name for this molecule.



A Pentanoic acid

B Hexanoic acid

C Heptanoic acid

D Octanoic acid

2 Which of the following is the correct IUPAC name for the semi-structural formula $CH_3CH_2CH_2CH_2CHCHCH_3$?

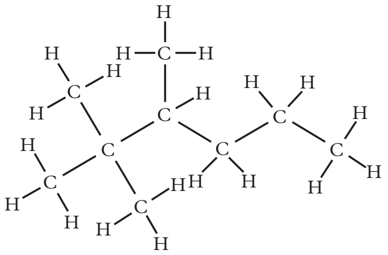
A Hept-5-ene

B Hept-2-ene

C Hex-2-ene

D Hex-5-ene

3 Consider the molecule below and the following statements:



I The name of this molecule is 2,2,3-trimethylhexane.

II It is an isomer of 2-methyloctane.

III It has four methyl groups.

IV It is a branched alkene.

Which of the above statements are correct?

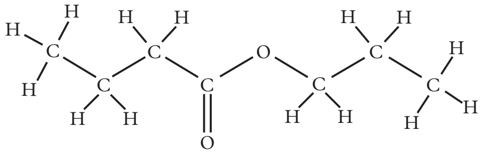
A I and II only

B I and IV only

C III and IV only

D II and III only

4 What is the correct structural name for this compound?



A Propyl butanoate

B Butyl propanoate

C Propyl propanoate

D Propyl pentanoate

5 How many structural isomers are there for C_4H_9F?

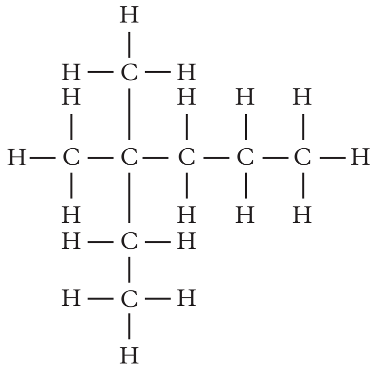
A 2

B 4

C 6

D 8

6 Identify the correct IUPAC systematic name of this molecule:



A 2-ethyl-2-methylpentane

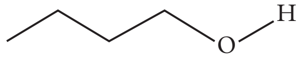
B 3,3-dimethylhexane

C 2-methyl-2-propylbutane

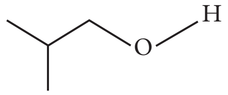
D 4,4-dimethylhexane

7 Identify which of the following is not a possible isomer of $C_4H_{10}O$:

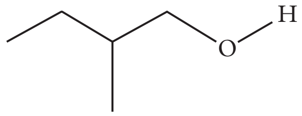
A



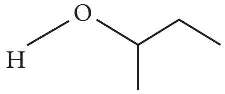
B



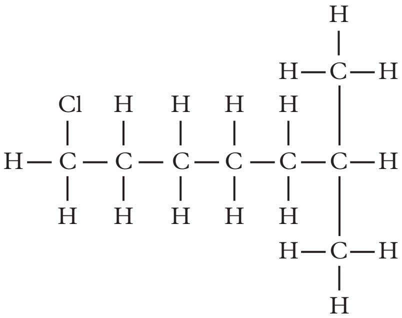
C



D



8 Identify the systematic name of the following molecule:



A 1-chloro-6,6-dimethylhexane

B 1–chloro–6–methylheptane

C 7–chloro–2–methylheptane

D 7-chloro-1,1-dimethylhexane

9 Viscosity decreases as temperature:

A increases.

B decreases.

C stays the same.

D none of the above.

10 Identify which of the following molecules will have the highest boiling point:

A Hexan-1-amine

B Hexan-3-amine

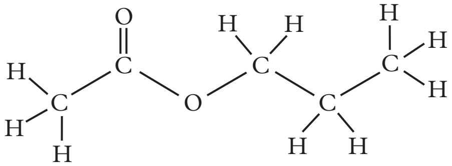
C Hexan-1-ol

D Hexan-3-ol

Short answer

Describe and explain

11 Consider the following compound:



a Identify the functional group in this compound.

b Give the systematic IUPAC name for this molecule.

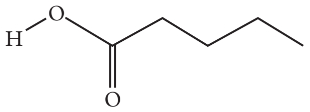
12 **a** Draw the structure of 3-chlorohex-2-ene.

b Explain whether there could be any structural isomers of this compound.

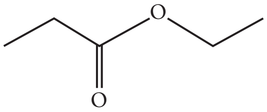
c Discuss how soluble 3-chlorohex-2-ene is and identify three of its isomers.

13 Consider the following compounds I–IV:

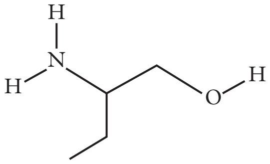
I



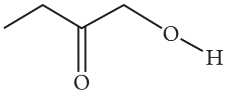
II



III



IV



a Identify and highlight all the functional groups present in the compounds.

b Name each of the functional groups present.

c Identify the systematic name of each compound.

14 In the following table, there are four compounds with similar molar masses:

Compound	Molar mass (g mol^{-1})	Boiling point (°C)
Pentane	72	36.1
Butanone	72	79.6
Butanol	74	117.7
Ethyl methanoate	74	?

a Draw the structural formula for each compound.

b Identify the intermolecular bonding between the compounds.

c Select a boiling point for ethyl methanoate from the suggestions below:

I Lower than pentane

II Between pentane and butanone

III Between butanone and butanol

IV Higher than butanol

d Explain why you selected your answer for part **c**.

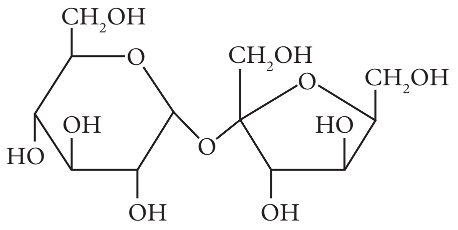
15 Identify whether a double bond is considered a functional group or not, and justify your answer.

16 A molecule has the empirical formula CH_2O and a molar mass of $60\,g\,mol^{-1}$.

a Identify the molecular formula of this molecule.

b Draw all the possible isomers of this molecular formula.

17 Maple syrup is primarily made of sucrose molecules shown below:



Sucrose

If you store maple syrup in the fridge, when you pour it on your pancakes, it is very thick and pours slowly. However, once it hits the warm pancakes, it becomes a lot thinner and flows more easily around the plate.

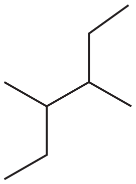
a Identify all the functional groups of the sucrose molecule.

b Identify the intramolecular bonding between sucrose molecules in maple syrup.

c Explain why maple syrup has a high viscosity.

d Explain what happens to the viscosity of maple syrup when it heats up.

18 Consider the molecule below:



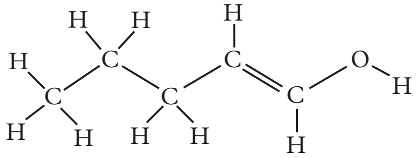
a Identify the longest main carbon chain in this compound.

b Identify the functional group(s) in this compound.

c Identify the systematic name of this compound.

d Write the semi-structural formula for this compound.

19 Explain why the name 1-hydroxypent-1-ene is not the correct name for this compound.



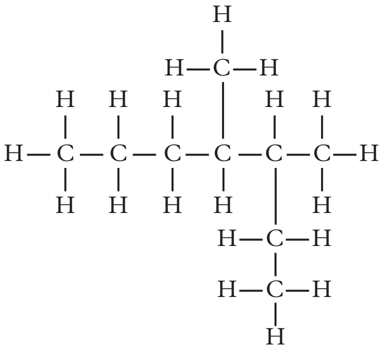
20 Describe the strength of all the bonds in the molecule ethanol.

21 Explain, with reference to electronegativity, why a C–F bond will be harder to break than a C–N bond.

22 Compare a primary alcohol and a secondary alcohol.

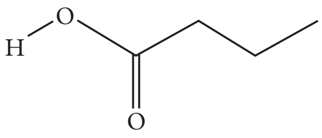
Apply and analyse

23 Apply IUPAC systematic naming to this molecule:



24 The boiling point of hexane is 69°C and the boiling point of hex-1-ene is 63°C. Compare the two compounds and explain the difference in their boiling points.

25 a Apply IUPAC systematic naming rules to these two compounds.



b Name the ester that would be formed when the two compounds are combined.

26 Draw all the possible structural isomers that contain a double bond and have the molecular formula C_5H_{10}. Apply the IUPAC naming rules to identify their names.

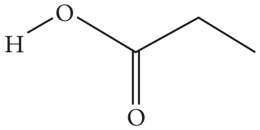
27 Apply the IUPAC systematic naming rules to the following semi-structural formulas:

$CH_3(CH_2)_3CHOHCH_3$
$CH_3CH(CH_3)CH_2CH_3$
$CH_3CH_2CHOHCH(CH_3)CH_2CH_3$
$CH_3CHCHCHNH_2CH_3$
$CH_3CH_3CHCH_2CHO$

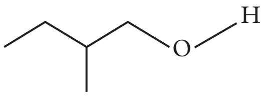
28 **a** Analyse the following compounds and rank their boiling points in order of lowest to highest.

b Apply IUPAC naming rules to each of the molecules.

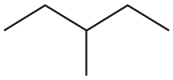
A



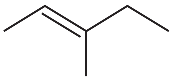
B



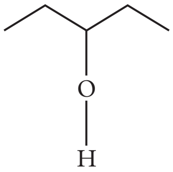
C



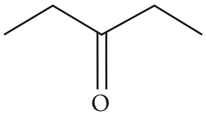
D



E



F



29 Discuss the polarity of each of the functional groups in Question 28 and identify if the molecules will be soluble in water.

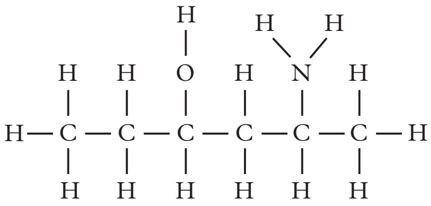
30 Consider a compound that contains carbon, nitrogen and hydrogen and has a molar mass of 71. Identify the molecular formula and draw all the possible isomers of this molecule, then apply the IUPAC naming rules to all the isomers to identify their correct systematic names.

31 Below are the semi-structural formulas of three organic compounds:

$CH_3CH_2COCH_2CH_3$
$CH_3CH_2CH_2CH_2COOH$
$CH_3C(CH_3)_2COOH$

The compounds have boiling points of 186°C, 163.7°C and 102°C. Compare the compounds' structures and determine which boiling point belongs to which compound.

32 Consider the following compound:



It could be named 2-aminohexan-4-ol or 5-aminohexan-3-ol.

Analyse the names and explain why both names could be considered correct. Then, decide which name is more correct considering the IUPAC rules. Justify your response.

33 A student makes the following statement:

'The longer the bond between two atoms, the less stable the molecule will be, so a C–I bond will be more stable than a C–H bond.'

Analyse the statement and decide if it is correct or not, then either expand the statement to make it clearer or rewrite it to make it correct.

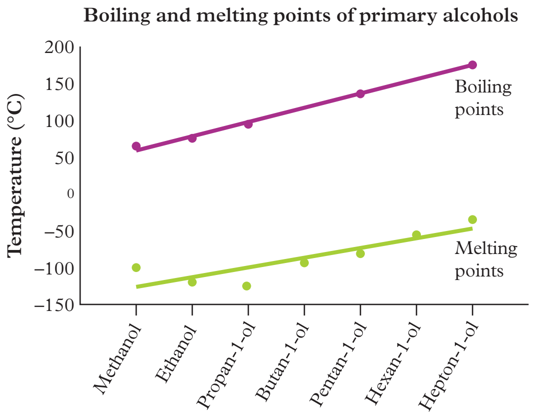
34 Consider the compounds CH_3CH_2OH and $CH_3CH_2CH_2CH_2CH_2CH_2CH_2COOH$.

a Apply IUPAC naming rules to these compounds.

b Name the ester that is formed when the two compounds are combined.

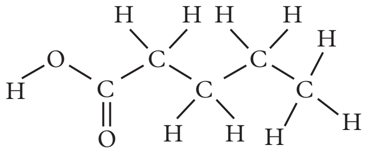
c Draw the ester.

35 Analyse the graph shown to answer the following questions.


Boiling and melting points of primary alcohols

- **a** Explain the trend in boiling points of primary alcohols.
- **b** Explain the trend in melting points of primary alcohols.
- **c** Explain what state pentan-1-ol will be at room temperature.
- **d** Predict the boiling points for butan-1-ol and hexan-1-ol from the graph.
- **e** Predict the melting point of octan-1-ol from the graph.

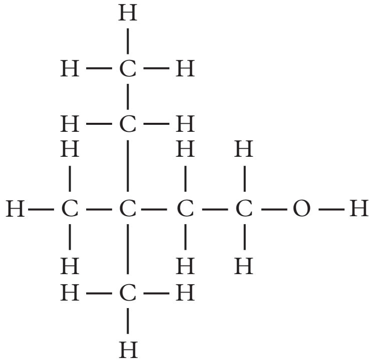
36 Analyse the following compound and identify the functional group. Explain why this functional group could only be on the end of the chain.



37 Compare the esters and explain the trends in the boiling points shown in the table.

Ester	Boiling point (°C)
Butyl ethanoate	126
Pentyl ethanoate	149
Hexyl ethanoate	155

38 Consider the compound below.



A student claims the compound is named 3-ethyl-3-methylbutan-1-ol, but this is incorrect. Apply the correct IUPAC naming rules to this structure and explain where the student went wrong in naming this compound.

39 Compare the following four semi-structural formulas:

$CH_2BrCHBrCH_2CH_3$

$CH_3CHBrCH_2CH_2Br$

$CH_3CHBrCHBrCH_3$

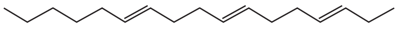
$CH_3CBr_2CH_2CH_3$

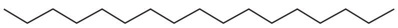
Describe what they have in common and apply IUPAC naming rules to each compound.

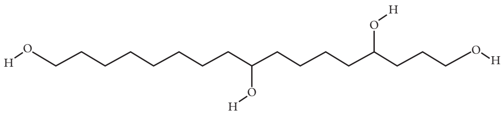
40 Compare the compounds methyl ethanoate and ethyl methanoate.

- **a** Draw the structural formulas for both compounds.
- **b** Discuss whether they are isomers.

41 Analyse the following skeletal structures and compare the viscosities of the three compounds:

A


B


C


Design and discuss

42 Discuss the relationship between:

- **a** electronegativity and bond strength
- **b** bond energy and bond stability.

43 You have been given two unknown organic compounds, X and Y. They have the same molecular formula, $C_4H_{10}O$, but different structural formulas.

- **a** Design a simple experiment to determine the boiling points of the two unknown organic compounds.
- **b** Identify the independent and dependent variables in your experiment.
- **c** Draw the structural formulas of the four possible isomers of $C_4H_{10}O$.
- **d** The boiling points of the isomers of $C_4H_{10}O$ are 117.7°C, 108°C, 99°C and 82°C. Discuss which isomer of $C_4H_{10}O$ matches each boiling point and explain your reasoning.
- **e** Discuss whether the experiment you designed in part **a** could determine the boiling points and distinguish the two unknown chemicals.

44 Evaluate the names of the following compounds and discuss the errors made in each, then correct the names:

- **a** 2-ethyl-3-methylbutane
- **b** 2-ethylpent-1-ene
- **c** 2-hydroxypentan-6-oic acid
- **d** 1-hydroxybutan-4-amine
- **e** 5-formylpentan-3-ol
- **f** 1,1-dimethyl-3-aminopropan-2-one
- **g** 2,2,3-chlorohexene
- **h** 2-aminopropane

45 There are three unbranched ester isomers with the molecular formula $C_4H_8O_2$ and boiling points of 77.1°C, 79.8°C and 82.0°C.

- **a** Draw and name the three unbranched esters.
- **b** Identify which boiling point matches each ester.
- **c** Discuss why the three esters have different boiling points even though they have the same molecular formula.

46 Discuss the difference in possible structures between a primary alcohol and a primary amine.

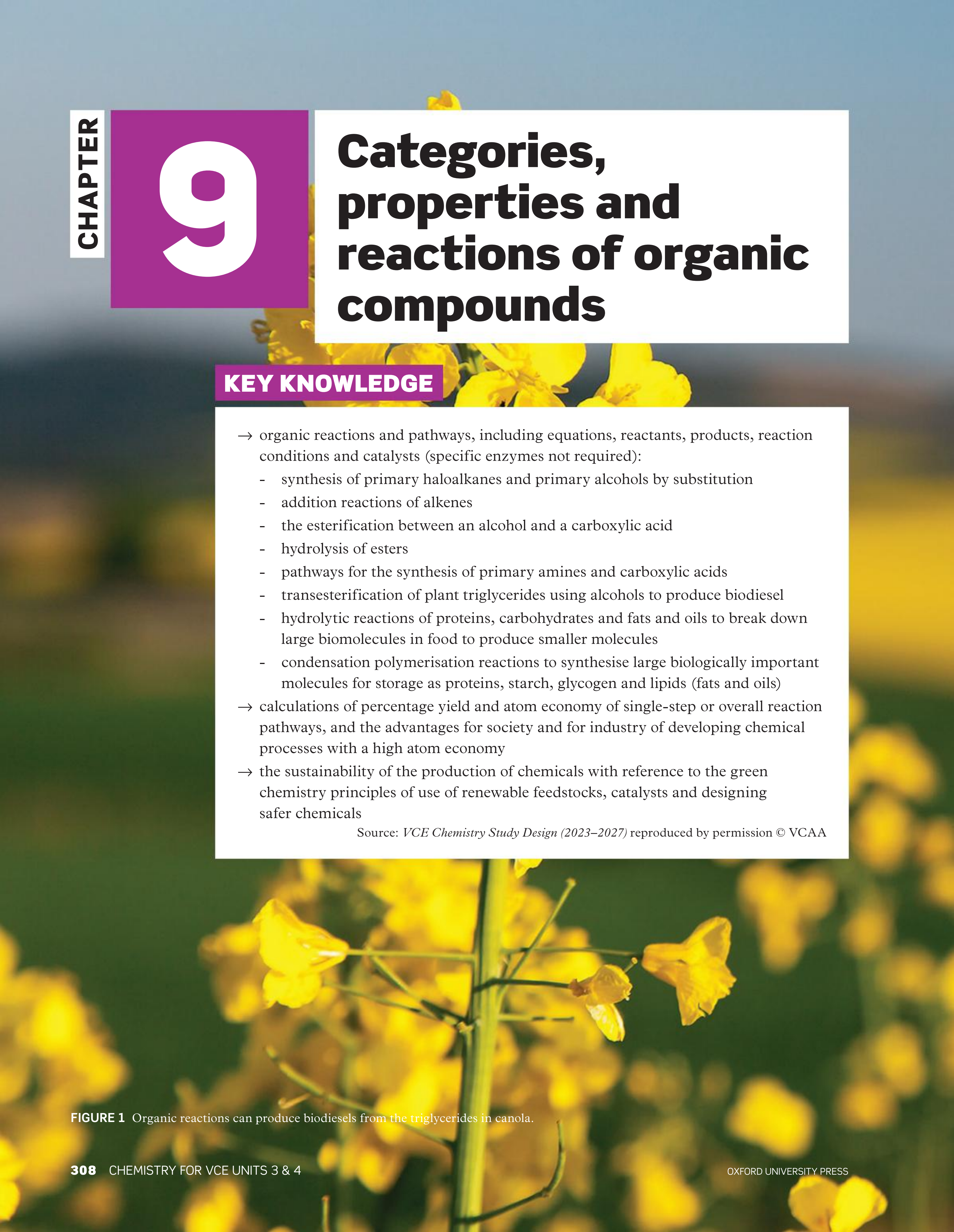
47 Consider the four types of structural representations that are used for organic compounds. Discuss the advantages and disadvantages of each of them and suggest a specific situation that each would be best used in.

You can find the following resources for this section in your obook pro:

pro

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.



CHAPTER 9

Categories, properties and reactions of organic compounds

KEY KNOWLEDGE

- → organic reactions and pathways, including equations, reactants, products, reaction conditions and catalysts (specific enzymes not required):
  - synthesis of primary haloalkanes and primary alcohols by substitution
  - addition reactions of alkenes
  - the esterification between an alcohol and a carboxylic acid
  - hydrolysis of esters
  - pathways for the synthesis of primary amines and carboxylic acids
  - transesterification of plant triglycerides using alcohols to produce biodiesel
  - hydrolytic reactions of proteins, carbohydrates and fats and oils to break down large biomolecules in food to produce smaller molecules
  - condensation polymerisation reactions to synthesise large biologically important molecules for storage as proteins, starch, glycogen and lipids (fats and oils)
- → calculations of percentage yield and atom economy of single-step or overall reaction pathways, and the advantages for society and for industry of developing chemical processes with a high atom economy
- → the sustainability of the production of chemicals with reference to the green chemistry principles of use of renewable feedstocks, catalysts and designing safer chemicals

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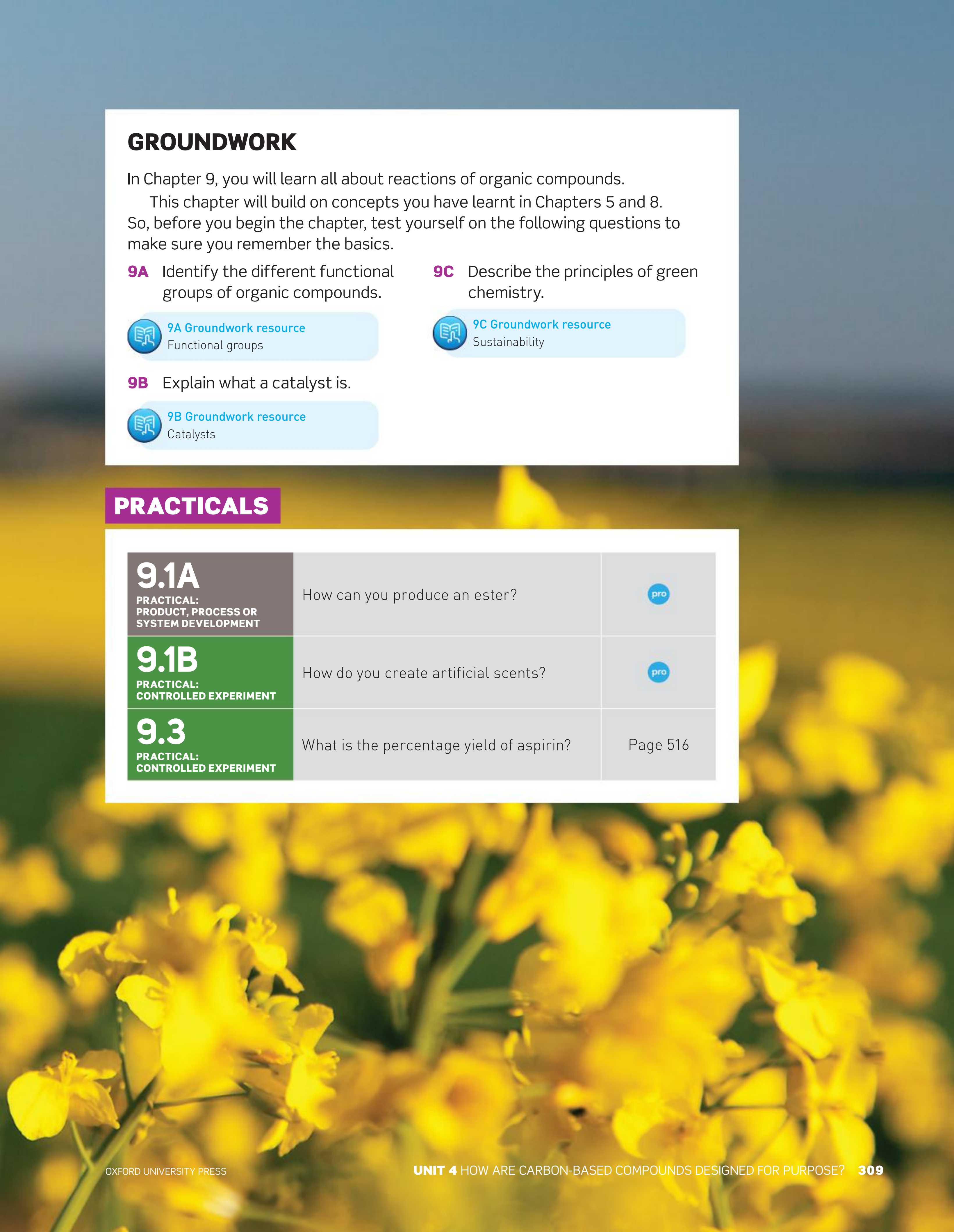
FIGURE 1 Organic reactions can produce biodiesels from the triglycerides in canola.



GROUNDWORK

In Chapter 9, you will learn all about reactions of organic compounds.

This chapter will build on concepts you have learnt in Chapters 5 and 8. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

9A Identify the different functional groups of organic compounds.

9A Groundwork resource
Functional groups

9B Explain what a catalyst is.

9B Groundwork resource
Catalysts

9C Describe the principles of green chemistry.

9C Groundwork resource
Sustainability

PRACTICALS

9.1A **PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT**	How can you produce an ester?	pro
9.1B **PRACTICAL: CONTROLLED EXPERIMENT**	How do you create artificial scents?	pro
9.3 **PRACTICAL: CONTROLLED EXPERIMENT**	What is the percentage yield of aspirin?	Page 516

9.1 Organic reactions

KEY IDEAS

In this topic, you will learn that:

- organic compounds can form new compounds through substitution, addition, oxidation, condensation and hydrolysis reactions
- some reactions require catalysts and specific conditions to occur
- an organic reaction pathway is created to synthesise a more complex molecule from smaller cheaper starting molecules; a pathway has multiple steps.

In previous chapters, we have looked at different functional groups of organic molecules. In this chapter, we will look at the reactions that organic molecules can undergo to form other molecules. We will also look at the conditions, reagents and catalysts that are required for these reactions to take place.

Synthesis of primary haloalkanes and primary alcohols by substitution

substitution reaction
a reaction where an atom in one molecule is replaced by an atom from another molecule; if the reaction starts with two reactants, it will end with two new products

Substitution reactions are reactions involving two reactants that are organic compounds. An atom or functional group from one reactant is replaced by a different atom or functional group from the other reactant. This results in the formation of two new substances.

Substitution reactions of alkanes

Alkanes can undergo substitution reactions with halogens such as chlorine and fluorine, where one of the hydrogen atoms is replaced by or substituted with a halogen atom. This type of reaction is used to synthesise haloalkanes and is shown in Figure 1.

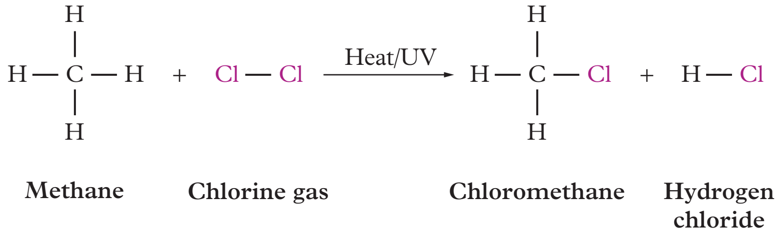


FIGURE 1 The substitution reaction between methane and chlorine to form chloromethane and hydrogen chloride

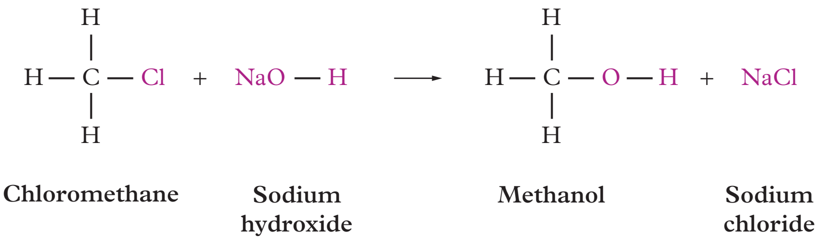
The substitution reaction in Figure 1 shows how one atom of chlorine replaces one of the hydrogen atoms to form the haloalkane chloromethane. The remaining chlorine and hydrogen combine to form the by-product hydrogen chloride. This reaction happens in the presence of UV light or heat.

Substitution reactions of haloalkanes

Haloalkanes can undergo substitution reactions to form alcohols and amines. In Figure 2, chloromethane can be reacted with water or sodium hydroxide to produce an alcohol.

9.1A
PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT
How can you produce an ester?
Go to your **Student obook pro**.

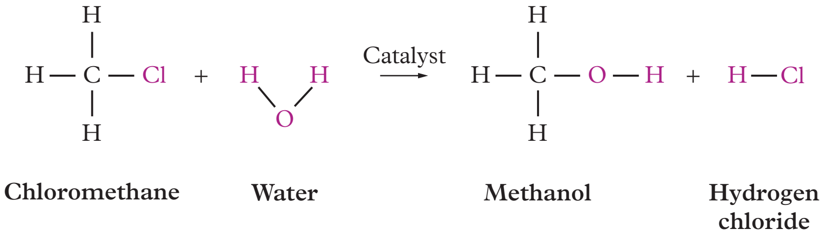
9.1B
PRACTICAL: CONTROLLED EXPERIMENT
How do you create artificial scents?
Go to your **Student obook pro**.



$$\mathrm{H{-}\overset{H}{\underset{H}{C}}{-}Cl + NaO{-}H \longrightarrow H{-}\overset{H}{\underset{H}{C}}{-}O{-}H + NaCl}$$

FIGURE 2 Chloromethane reacts with sodium hydroxide to form methanol.

In Figure 2, the negatively charged hydroxide group from the sodium hydroxide molecule replaces the electronegative chlorine atom on the haloalkane. This forms an alcohol (methanol) and sodium chloride.



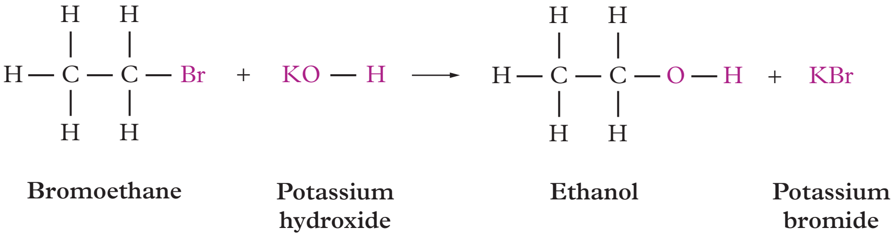
$$\mathrm{H{-}\overset{H}{\underset{H}{C}}{-}Cl + H_2O \xrightarrow{Catalyst} H{-}\overset{H}{\underset{H}{C}}{-}O{-}H + H{-}Cl}$$

FIGURE 3 Chloromethane reacts with water to form methanol and hydrogen chloride.

Haloalkanes can also undergo substitution reactions with water to form an alcohol. Figure 3 shows the substitution reaction between chloromethane and a water molecule. This reaction has to happen in the presence of a catalyst, as it is less likely to proceed than the substitution reaction with a hydroxide ion. This is because water lacks the full negative charge that the hydroxide ion has.

A substitution reaction with a haloalkane to form an alcohol does not always use chloroalkanes, sodium hydroxide or water as reactants. Any of the halogens can participate in this reaction and potassium hydroxide can also be a reactant. In Figure 5, you can see the reaction between potassium hydroxide and bromoethane to form ethanol and potassium bromide.

FIGURE 4 Methanol used in wash bottles can be sythesised by substitution reactions.



$$\mathrm{H{-}\overset{H}{\underset{H}{C}}{-}\overset{H}{\underset{H}{C}}{-}Br + KO{-}H \longrightarrow H{-}\overset{H}{\underset{H}{C}}{-}\overset{H}{\underset{H}{C}}{-}O{-}H + KBr}$$

FIGURE 5 The substitution reaction to form ethanol from bromoethane and potassium bromide

Both of these reactions show haloalkanes reacting to form primary alcohol**s** such as methanol and ethanol.

Substitution reactions can take place with any of the haloalkanes and with any halogen positioned along the chain. This is shown in Figure 6, where 2-bromobutane reacts with sodium hydroxide to produce butan-2-ol, a secondary alcohol. The hydroxide will always take the position of the halogen in a substitution reaction.

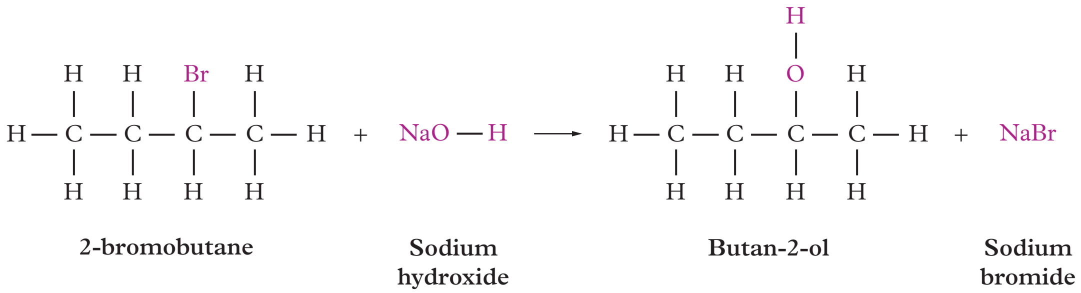


FIGURE 6 The substitution reaction of 2-bromobutane and sodium hydroxide to produce butan-2-ol

Addition reactions of alkenes

addition reaction
a reaction in which two molecules add together to form one new molecule; usually involving a double bond

In **addition reactions**, two or more reactants combine to form one single product. In this chapter we will look at the addition reactions of alkenes to form alkanes, haloalkanes, alcohols and alkane polymers.

Addition reactions to form alkanes

hydrogenation
the addition of H_2 to an alkene to form an alkane

Reacting an alkene with hydrogen gas in the presence of a nickel catalyst will form an alkane. This is referred to as **hydrogenation**. An example of this addition reaction is shown in Figure 7, with ethene and hydrogen gas reacting to form ethane. In this reaction, the alkene double bond breaks and the two hydrogen atoms attach where the double bond once was, forming an alkane.

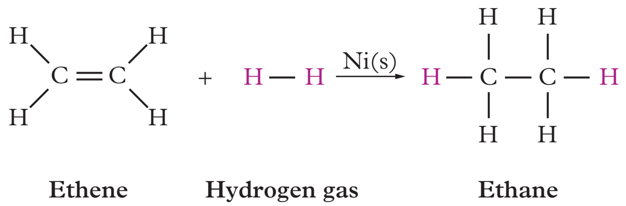


FIGURE 7 The addition reaction between ethene and hydrogen gas to form ethane

Addition reactions to form haloalkanes

hydrogen–halide
a molecule containing a hydrogen and a halogen; can be H–Br, H–Cl, H–F or H–I

Alkenes can also undergo addition reactions with **hydrogen–halides** (H–Br, H–F, H–Cl, H–I) and halogen gases (Br_2, F_2, Cl_2, I_2) to form haloalkanes. The hydrogens and halogens will always attach to the carbon atoms where they were once double-bonded. You can see this in Figures 8 and 9, where the additions are in purple.

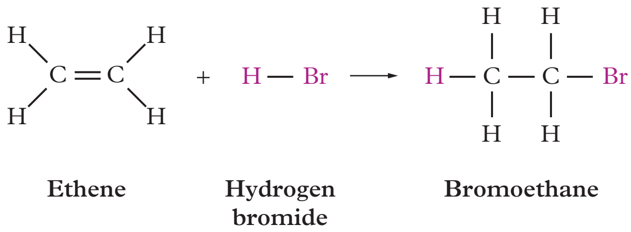


FIGURE 8 Ethene and hydrogen bromide react together in an addition reaction to form bromoethane.

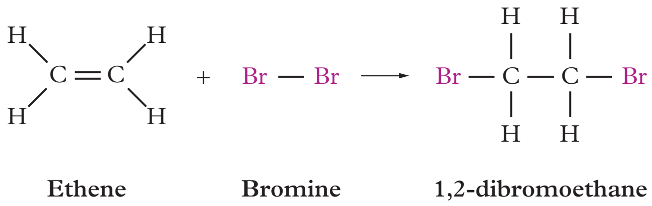


FIGURE 9 Ethene and bromine gas react together in an addition reaction to form 1,2-dibromoethane.

Addition reactions to check for carbon–carbon double bonds

Bromine is often used to check if organic compounds are unsaturated. Bromine is a brown/orange coloured liquid. If an organic compound is saturated, the colour of the solution will remain brown. If it is unsaturated, the double bond and bromine will react to form an alkane with two bromines attached to the carbon atoms where the double bond once was. This forms a clear compound.

Study tip

Be careful when you are drawing structures of an addition reaction as in Figure 9. Make sure you draw the new additions on the carbon atoms where the double bond was – one on each side of the old carbon–carbon double bond.

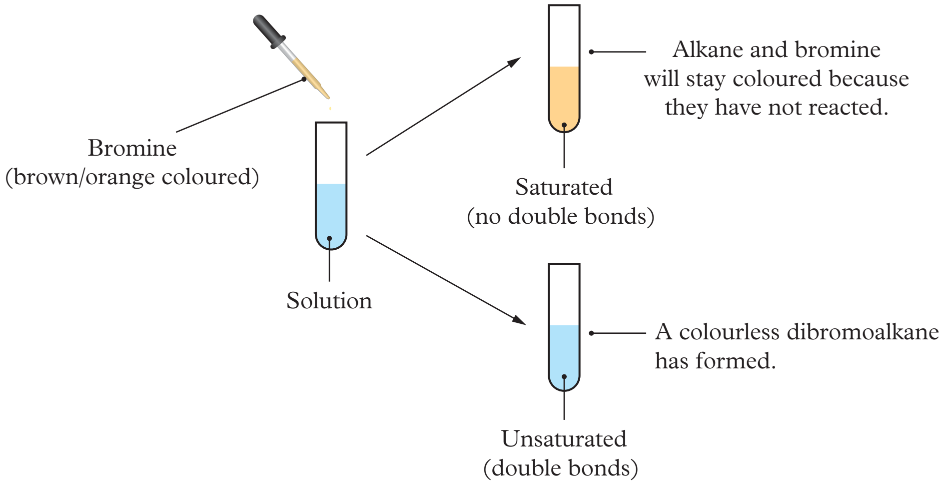


FIGURE 10 Bromine is used to check if an organic compound is saturated or unsaturated.

Addition reactions to form alcohols

Alkenes can also react in an addition reaction with water to form an alcohol. This is referred to as hydration of alkenes and must be conducted at high temperatures (300°C) with a solid phosphoric acid (H_3PO_4) catalyst.

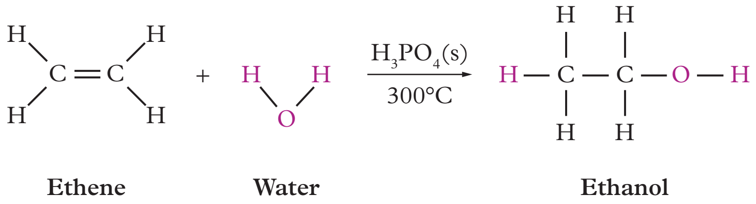


FIGURE 11 The addition reaction of water and ethene to form ethanol

When longer chain alkenes undergo addition reactions, two possible products can form. For example, water can be added to pent-2-ene to form pentan-2-ol or pentan-3-ol (Figure 12).

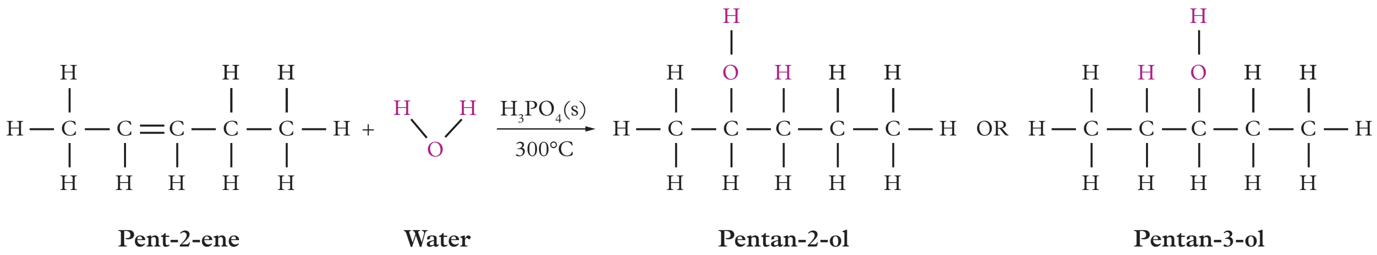


FIGURE 12 The addition reaction of water and pent-2-ene to form pentan-2-ol and pentan-3-ol

Addition reactions to form polymers

In Unit 1, you learnt that alkenes can react in addition reactions to form polymers, in a process known as addition polymerisation. The alkene is the monomer and it combines together in repeating units to form long polymer chains.

Figure 14 shows the monomer ethene undergoing addition polymerisation to form the polymer polyethylene. Polyethylene is a polymer commonly used to make cling film for kitchen and grocery bags.

FIGURE 13 Some plastic bags are made of polyethylene.

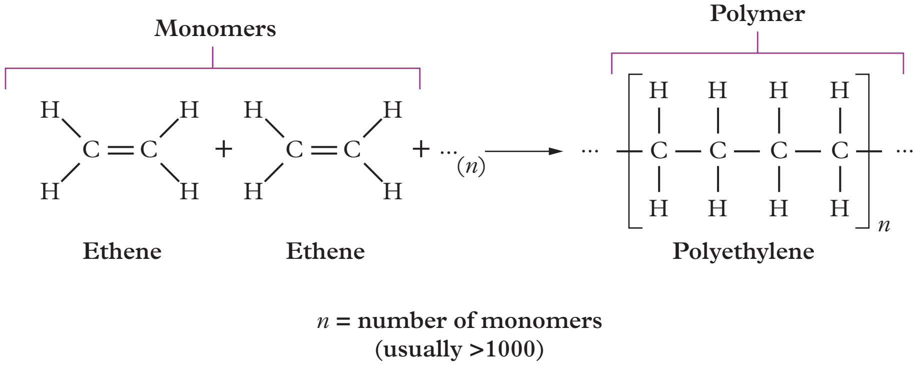


FIGURE 14 The addition reaction of the monomer ethene to form the polymer polyethylene

Pathways for the synthesis of primary amines and carboxylic acids

Synthesis of primary amines

Earlier in the topic, we looked at substitution reactions of haloalkanes. Haloalkanes can also undergo a substitution reaction with ammonia (NH_3) to form primary amines. This substitution reaction is shown in Figure 15. The ammonia breaks into NH_2 and hydrogen, NH_2 replaces the Cl on chloromethane to form methanamine and the free H and Cl combine to form hydrogen chloride.

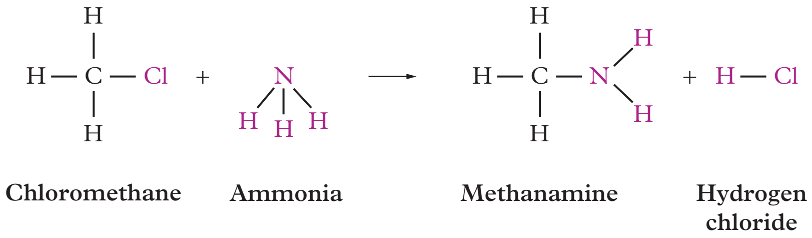


FIGURE 15 Chloromethane reacts with ammonia to form methanamine and hydrogen chloride.

Amines can also be formed from alkenes, but you cannot just attach an NH_2 group at the double bond. For an amine to form from an alkene, the reaction requires an intermediary step – the addition of a halogen – to then undergo a substitution reaction with ammonia.

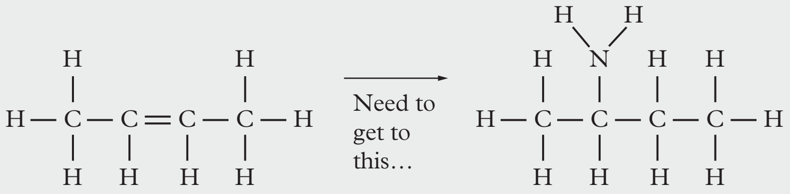
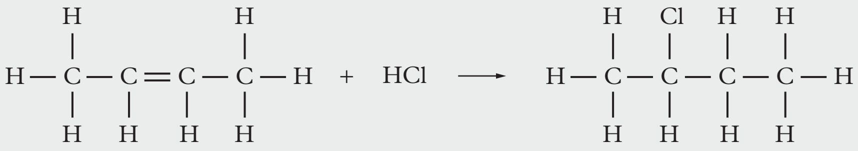
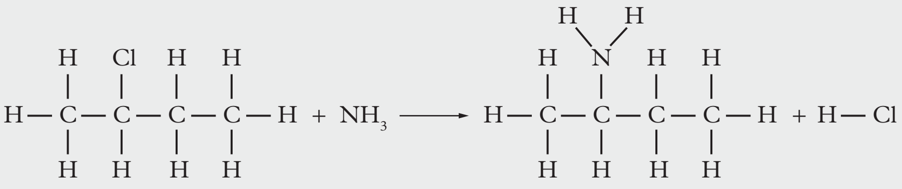
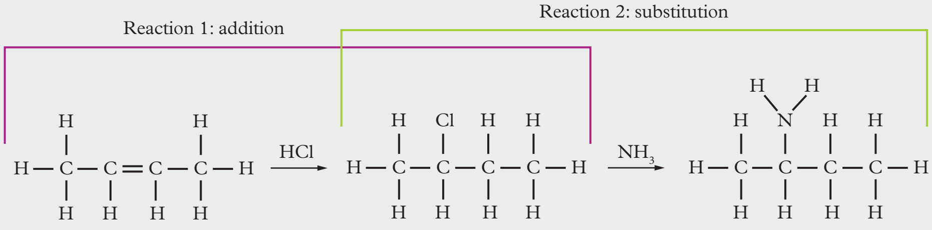
You can see an example of an alkene being synthesised into a primary amine in Worked example 9.1A.

FIGURE 16 Primary amines are commonly used in azo dyes, which contain an azo bond N=N.

9.1 WORKED EXAMPLE

IDENTIFYING PATHWAYS FOR SYNTHESISING AN AMINE

Create a pathway to synthesise butan-2-amine from but-2-ene.

Think	Do
Step 1: Consider your starting product and the final product and assess whether you can get from start to finish in one step.	The starting product is but-2-ene and the final product is butan-2-amine.  Need to get to this… **But-2-ene** **Butan-2-amine** You cannot get from but-2-ene to butan-2-amine in one step, so the pathway will require multiple steps.
Step 2: Think about what is needed in order to synthesise an amine. You need to have a haloalkane and complete a substitution reaction, so you first have to turn but-2-ene into a haloalkane. You can do this via an addition reaction using a hydrogen halide.	 **But-2-ene** **2-chlorobutane**
Step 3: Now that you have 2-chlorobutane, complete a substitution reaction with ammonia to produce butan-2-amine.	 **2-chlorobutane** **Butan-2-amine**
Step 4: Combine the two reactions into one complete pathway and write your final answer.	 Reaction 1: addition Reaction 2: substitution **But-2-ene** **2-chlorobutane** **Butan-2-amine**

Study tip

When completing oxidation reactions of alcohols, there are two main oxidising agents you can use. You should be familiar with $KMnO_4/H^+$ (acidified potassium permanganate) and $K_2Cr_2O_7/H^+$ (acidified potassium dichromate).

You can also show them without spectator ions (MnO_4^- and $Cr_2O_7^{2-}$). When you are writing your own oxidation reactions, stick with the one oxidising agent you remember the best and *always* remember to include H^+.

Synthesis of carboxylic acids

The pathway to synthesising a carboxylic acid from an alkane or alkene is a multi-step reaction pathway. To understand this pathway, you will need to understand the oxidation of a primary alcohol into a carboxylic acid.

Oxidation of primary alcohols

Primary alcohols can be oxidised to form carboxylic acids. This can happen naturally over time with oxygen (O_2) in the air. The hydroxyl group can also be oxidised to form a carboxyl group using a strong acidified oxidising agent such as potassium permanganate ($KMnO_4$) or potassium dichromate ($K_2Cr_2O_7$).

The oxidation of a primary alcohol to form a carboxylic acid has an intermediary step at which an aldehyde is formed. This step occurs when there is not enough heat or time for the full oxidation of into a carboxylic acid, as shown in Figure 17.

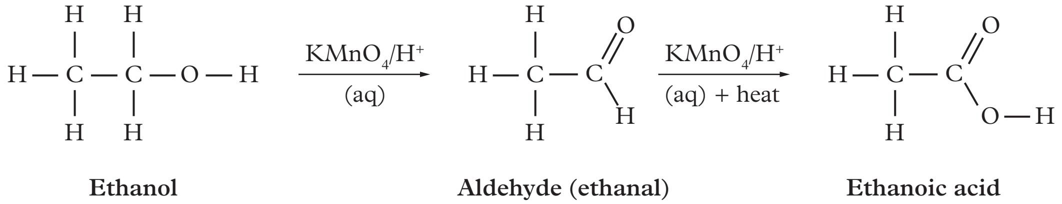


FIGURE 17 The oxidation reaction of ethanol to ethanal and then to ethanoic acid

Longer reaction times and higher temperatures will favour carboxylic acids being formed. On the other hand, milder reaction conditions, such as shorter reaction times and lower temperatures, will favour the formation of an aldehyde.

Oxidation of secondary alcohols

Secondary alcohols can also be oxidised, but they do not form carboxylic acids or aldehydes, as primary alcohols do. Due to the placement of the hydroxyl group on a secondary alcohol, they will oxidise to form ketones. If you compare Figures 17 and 18, you will see that the oxidation process is the same, but a secondary alcohol becomes a ketone (a double-bonded oxygen attached to a carbon along the chain).

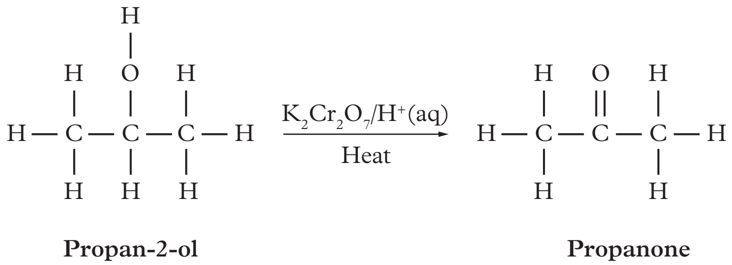


FIGURE 18 A secondary alcohol oxidises to form a ketone.

Tertiary alcohols do not normally oxidise in the presence of an oxidising agent. This is because the carbon with the hydroxyl group is attached to multiple alkyl groups, and does not have enough bonding pairs to oxidise OH into =O.

Reaction pathway to form a carboxylic acid

Now that you understand how the oxidation of an alcohol can form carboxylic acid, you can look at the reaction pathways that alkanes or alkenes undergo to form carboxylic acids.

Alkanes use a three-step reaction to form a carboxylic acid. In Figure 19, you can see how propanoic acid is formed in three steps from the alkane propane.

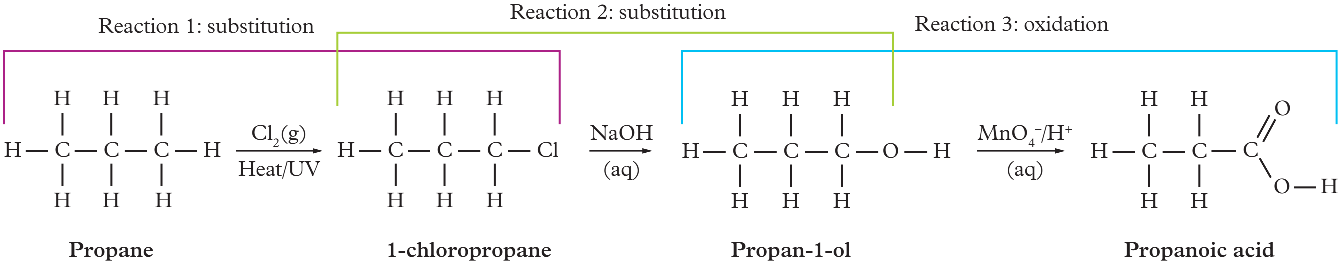


FIGURE 19 The reaction pathways containing two substitution reactions and one oxidation reaction to produce propanoic acid from propane

You may have noticed that the by-products formed in reaction pathways (such as HCl and NaCl produced by the first two substitution reactions from Figure 19) are not always shown. You should be conscious that these products are formed (as this is helpful when calculating atom economy) but you do not need to include them in the pathway if they are not used in the following steps.

Alkenes use a two-step reaction to form a carboxylic acid. As we learnt earlier in the topic, an alkene can react with water and a phosphoric acid catalyst to form an alcohol. This addition reaction is the first step in the reaction pathway to forming a carboxylic acid. The alcohol formed from this step can then be oxidised into a carboxylic acid. An example of this reaction pathway is shown in Figure 20.

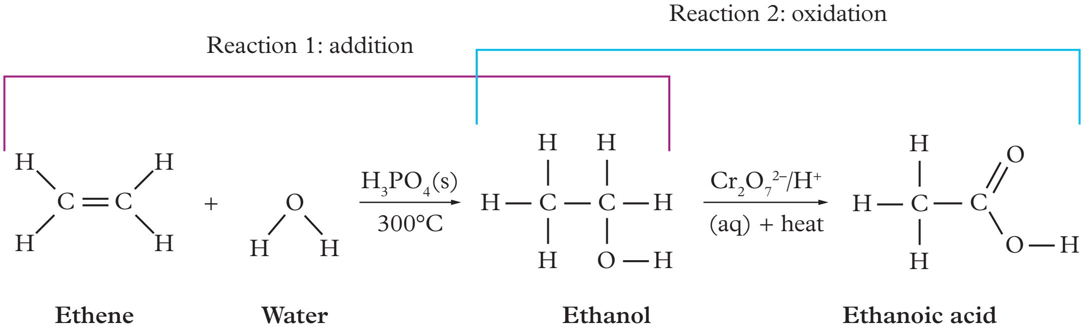


FIGURE 20 The reaction pathway containing addition and oxidation reactions to form ethanoic acid from ethene

Figures 19 and 20 are just two examples of the different pathways that could be completed to form a carboxylic acid. Both these examples can be used with any starting alkane or alkene.

Study tip

Make a table or chart of all the reactions that you learnt in this chapter. This will help you a lot when you are completing reaction pathways so you can quickly refer to it when revising.

Study tip

When completing reaction pathways, just take it one step at a time. Look for hints in the reactants, products or catalysts to help you figure out what reaction is required.

Esterification reactions

In organic chemistry, a **condensation reaction** occurs when two molecules are added together to form a larger single molecule and water (or another small molecule) as a by-product. Specifically, the condensation reaction between a carboxylic acid and an alcohol is also called **esterification**. This is a type of condensation reaction where an ester is formed.

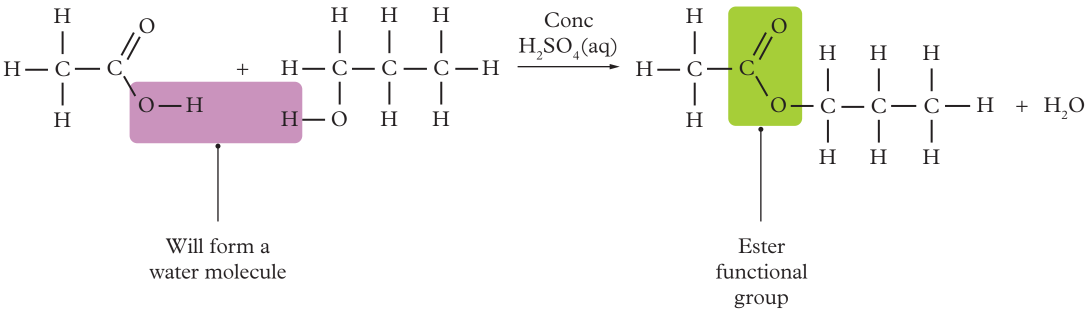


FIGURE 21 The condensation reaction between ethanoic acid and propanol to form the ester propyl ethanoate

condensation reaction
a reaction in which two molecules are added together to form a new larger molecule and a water molecule, or another small molecule

esterification
a condensation reaction to form an ester and a water molecule from an alcohol and carboxylic acid

An example of esterification is shown in Figure 21 on the previous page, where ethanoic acid reacts with propanol to form propyl ethanoate. The hydroxyl group from the carboxylic acid and a hydrogen from the hydroxyl group on the alcohol react to form a water molecule. The remaining carbon and oxygen atoms react to form the ester.

The esterification reaction is reversible, meaning that the ester can reform the carboxylic acid and alcohol reactants. To make sure this reaction proceeds in one direction to form the ester, concentrated sulfuric acid (H_2SO_4) is required as a catalyst.

Producing esters from alkanes and alkenes

An ester can be synthesised using alkanes or alkenes as raw materials. Alcohols and carboxylic acids derived from these compounds can then be combined in a condensation reaction to form an ester.

Let's consider the ester methyl propanoate. To form this ester, you need methanol and propanoic acid, which means that you would have to start with methane and propane or propene. The steps to form methyl propanoate from methane and propene are as follows:

1. Starting with methane, first complete the substitution reaction with chlorine to form chloromethane. Then, complete a second substitution reaction with sodium hydroxide to produce methanol.

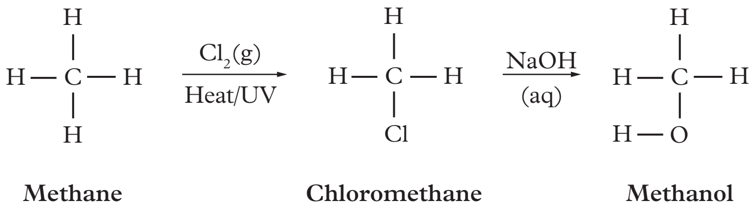


FIGURE 22 The pathway for synthesising methane into methanol

2. Next, you need to form propanoic acid. You could start with propane, but this example uses propene. The first step is the addition of water to propene, or hydration of propene, to form propan-1-ol. Propan-2-ol may also be formed in this reaction, and will need to be separated out. The second step is the oxidation of propan-1-ol into propanoic acid.

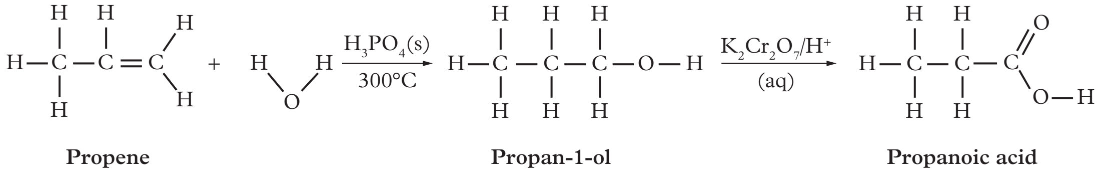


FIGURE 23 The pathway for synthesising propene into propanoic acid

3. Now that you have propanoic acid and methanol, you can combine them in a condensation reaction. This reaction uses concentrated sulfuric acid as a catalyst to form the ester methyl propanoate.

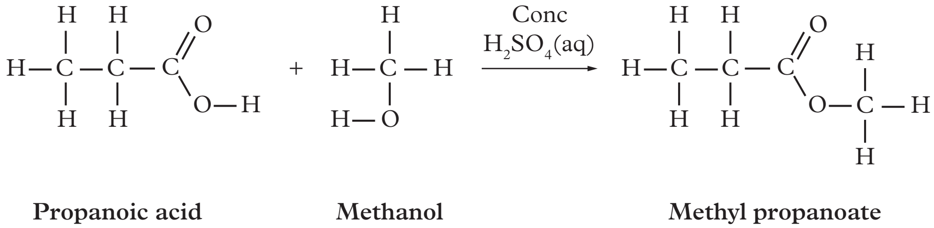


FIGURE 24 The condensation reaction between methanol and propanoic acid

The full reaction pathway for the reaction to synthesise methyl propanoate from methane and propene is shown in Figure 25.

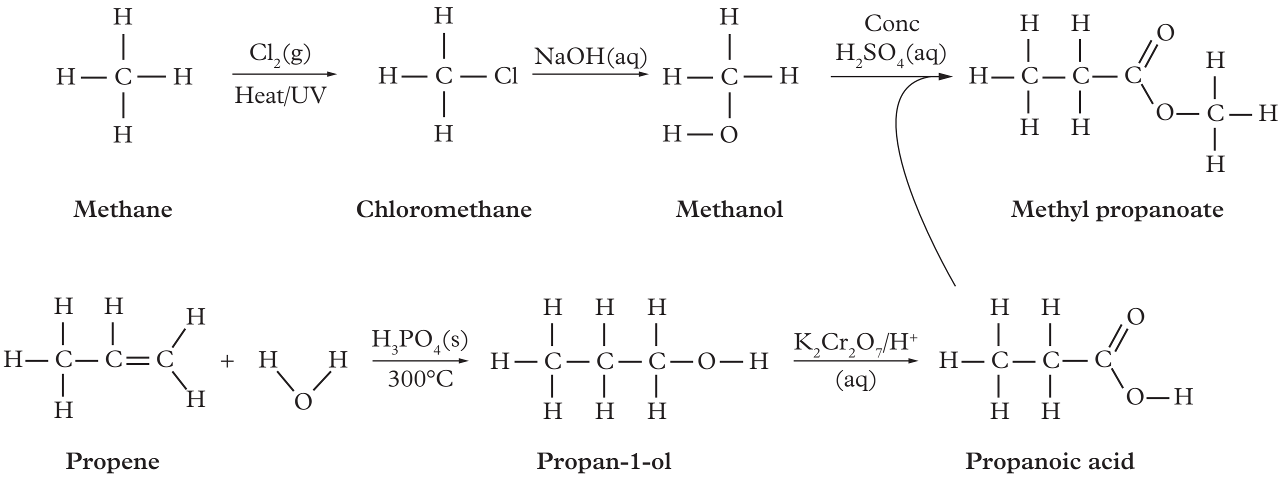


FIGURE 25 The reaction pathway for the synthesis of methyl propanoate from propene and methane

This is just one of the pathways that could be used to synthesise methyl propanoate, depending on the available starting products and desired number of steps. Have a go at designing a pathway to synthesise an ester in Challenge 9.1.

9.1 Challenge
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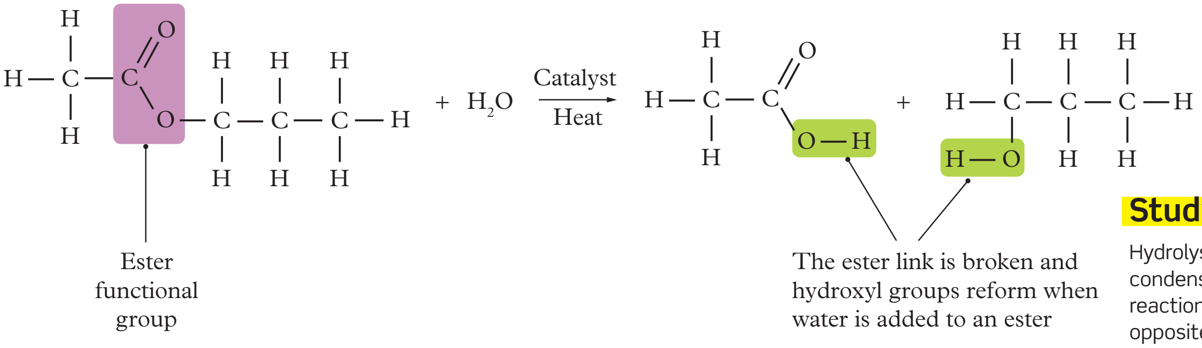
Hydrolysis of esters

Hydrolysis reactions are reactions in which a water molecule breaks the bonds in another molecule to form new products. A water molecule is added in the presence of heat and a catalyst to break the ester bond to form the alcohol and the carboxylic acid. For esters, it is the reverse reaction of the condensation reaction used to form the ester. In Figure 26, you can see propyl ethanoate hydrolysing back into propanol and ethanoic acid.

hydrolysis reaction the reverse of a condensation reaction – water is added to a reaction to form two or more new molecules



The ester link is broken and hydroxyl groups reform when water is added to an ester

FIGURE 26 The hydrolysis of propyl ethanoate back into propanol and ethanoic acid

Study tip

Hydrolysis and condensation reactions are opposites!

The hydrolysis of esters is usually catalysed by a strong acid catalyst, but can sometimes be completed in an alkaline solution. This is the process used for making soap from triglycerides, which you can explore in Skill drill 9.1 on the next page.

FIGURE 27 Soap can be produced by the hydrolysis of triglycerides.

9.1 SKILL DRILL

Evaluating an experiment to determine which oil makes the best soap

Key science skill: Analyse and investigate data and investigation methods

Soap is made using a hydrolysis reaction between a triglyceride molecule and a strong base (Figure 28). Three molecules of sodium hydroxide (the base) react with the three ester bonds in the triglyceride to form a molecule of glycerol.

Two students decided to investigate what oil would make the best soap. They reacted three different oils (coconut, vegetable and olive oil) with NaOH using the method below for each soap. They decided that the best soap would be the one that set the hardest after three days, smelt nice and was the softest when washing.

Materials:

- 500 mL of oil
- 150 mL of 1 M sodium hydroxide
- Hot plate
- Clean beakers
- Thermometer
- Whisk or stirring rod
- Trays to set soap in

Method:

1. In a clean beaker, add the oil and sodium hydroxide.
2. Place the beaker on the hot plate and stir the mixture until it is heated to around 60°C.
3. Carefully remove the beaker from the hot plate and continue stirring rapidly (but carefully) until the mixture thickens.
4. Once the mixture has thickened and changed colour, pour it into the tray to set.
5. After three days, the soap should be set and can be removed from the tray for testing.

Results:

Oil	Set hardness of soap	Smell	Gentleness of soap when washing hands
Vegetable	Set hard	Worst smell	Not softening, but not drying
Olive	Not set well; very soft and slippery	Smelled fine, but not great	Very softening on the hands
Coconut	Set hard	Nicest smell	Very drying

Conclusion: After testing the soaps, the students decided that none of the oils makes the best soap.

Practise your skills

1. Identify the independent and dependent variables in this investigation.
2. Explain what makes this reaction a hydrolysis reaction.
3. Design a step for the method that would make the students' results more reliable.
4. Analyse the results and explain the students' conclusion.
5. Propose a question that could follow on from this investigation.

Need help analysing and evaluating data and investigation methods? See Topic 1.8 (page 24).

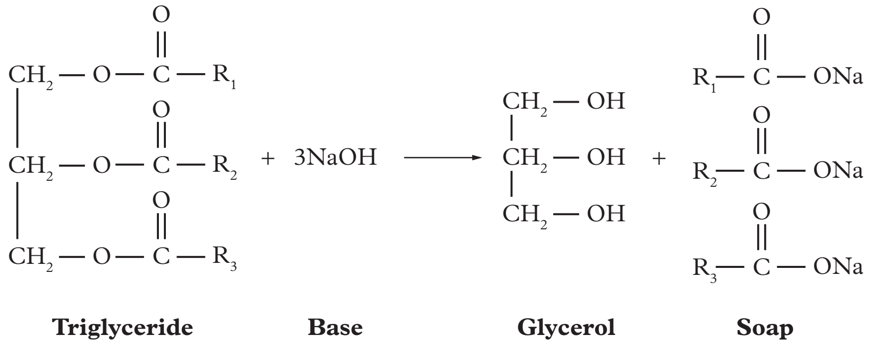


FIGURE 28 The reaction between a triglyceride (oil) and a base to form soap and glycerol

9.1 CHECK YOUR LEARNING

Describe and explain

1 Describe the following reactions.
  - a Addition
  - b Substitution
  - c Oxidation

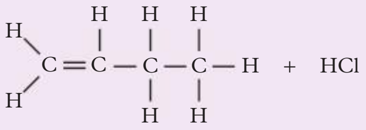
2 Identify the catalysts required for the following reactions.
  - a Butanol → butanoic acid
  - b Propane + Cl_2 → 1-chloropropane + HCl
  - c Hex-1-ene + H_2O → hexan-1-ol
  - d Methanoic acid + methanol → methyl methanoate
  - e Pentan-2-ol → pentan-2-one

3 Describe the types of reactions occurring in Question 2.

Apply, analyse and compare

4 Four incomplete reactions are shown on the right. Analyse the reactions and answer the following questions.
  - a Identify the types of reactions occurring.
  - b Identify any catalysts/conditions required.
  - c Complete the reactions.

I



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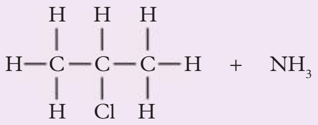
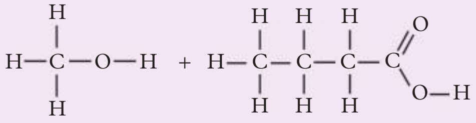


III



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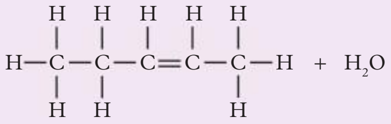
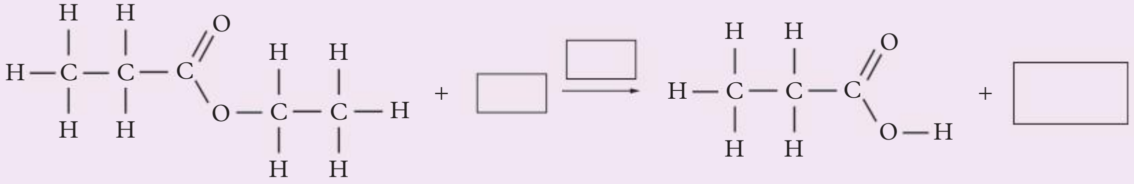
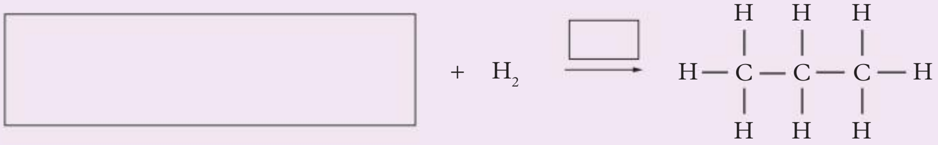


5 Analyse the following reactions and fill in the blanks.

a



b



Design and discuss

6 Propanol has two isomers.
  - a Draw the structural formulas of the two isomers of propanol.
  - b Design reactions to form the two isomers of propanol.
  - c Describe the difference between the products formed when the two isomers undergo oxidation.

7 Discuss the differences between a hydrolysis reaction and a condensation reaction.

8 Show the reaction pathways for when the following compounds are added to but-2-ene:
  - a Br_2
  - b H_2O
  - c HCl

9.2 Organic reactions of large biomolecules

KEY IDEAS

In this topic, you will learn that:

- carbohydrates, proteins, fats and oils are hydrolysed in our bodies to supply us with essential smaller molecules and energy
- small molecules within our bodies can undergo condensation polymerisation reactions to be stored as larger molecules
- biodiesel can be formed from the esterification reaction between a plant triglyceride and methanol.

Reactions of large biomolecules

biomolecule
a molecule produced by living cells and organisms

nucleic acid
a biomolecule that makes up the genetic material

peptide bonds
a $CONH_2$ bond; also known as an amide bond

secondary amide
an amide in which the nitrogen atom is attached to one hydrogen atom and one R group

The food we consume provides our bodies with the energy needed to complete the metabolic reactions that keep our bodies functioning properly. This nutrition comes in the form of **biomolecules**, which are produced by cells and living organisms. There are four types of biomolecules: proteins, carbohydrates, lipids and **nucleic acids**. Humans consume carbohydrates, proteins, fats and oils to maintain metabolic function.

In this topic, we will look at the hydrolysis reactions that break these large biomolecules into smaller molecules for transport and use in our bodies. We will also look at the condensation reactions used to reform them from their smaller molecules for storage.

Proteins

A protein is a long chain of specifically ordered amino acids linked together by amide or **peptide bonds**. In Chapter 8, you learnt about primary amides. A **secondary amide** has two alkyl groups attached to the carbonyl and amine (Figure 1).

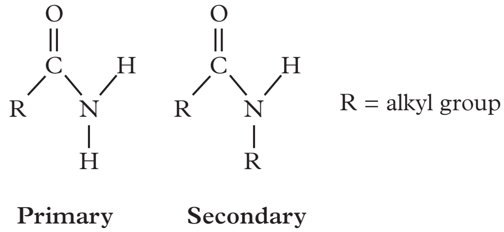


FIGURE 1 Primary and secondary amide functional groups

Primary amides occur at the ends of protein chains. Secondary amides occur within protein chains and act as the peptide bonds that link the amino acids together.

FIGURE 2 The human body relies on the consumption of proteins, fats and carbohydrates from food to maintain metabolic function.

Amino acids

Amino acids are the building blocks of proteins (Figure 3). There are 20 naturally occurring amino acids called the alpha amino (α-amino) acids. All α-amino acids have a similar backbone structure consisting of a central carbon that is surrounded by a hydrogen atom, carboxyl group, an amino group and an R group (Figure 4).

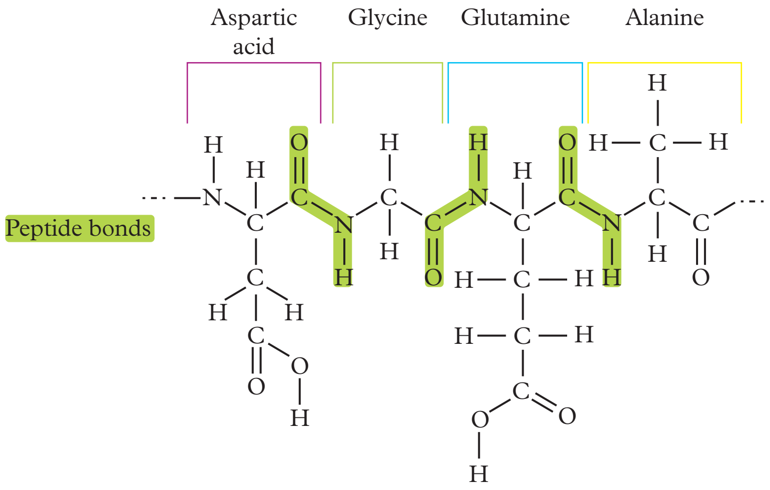


FIGURE 3 A protein is made of a sequence of amino acids connected by peptide or amide bonds.

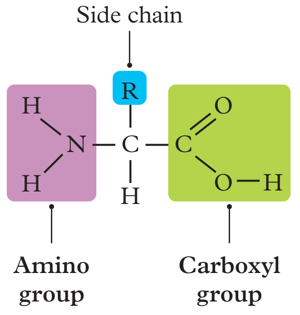


FIGURE 4 The general structure of an α-amino acid

R groups

The R group differs for each of the 20 α-amino acids. Each amino acid structure can be found in the VCAA data book. In Figure 5, you can see the four α-amino acids listed below, and their R groups in blue:

- Aspartic acid, with an R group of CH_2COOH
- Glycine, with an R group of just a single H atom
- Glutamine, with an R group of CH_2CH_2COOH
- Alanine, with an R group of CH_3

The individual α-amino acids can be separated by adding a water molecule to break each of the amide links within the protein.

Study tip

A peptide bond or link can also be known as an amide bond or link. The difference is that only amino acids form peptide bonds, but any amino group and carboxyl group can form an amide bond.

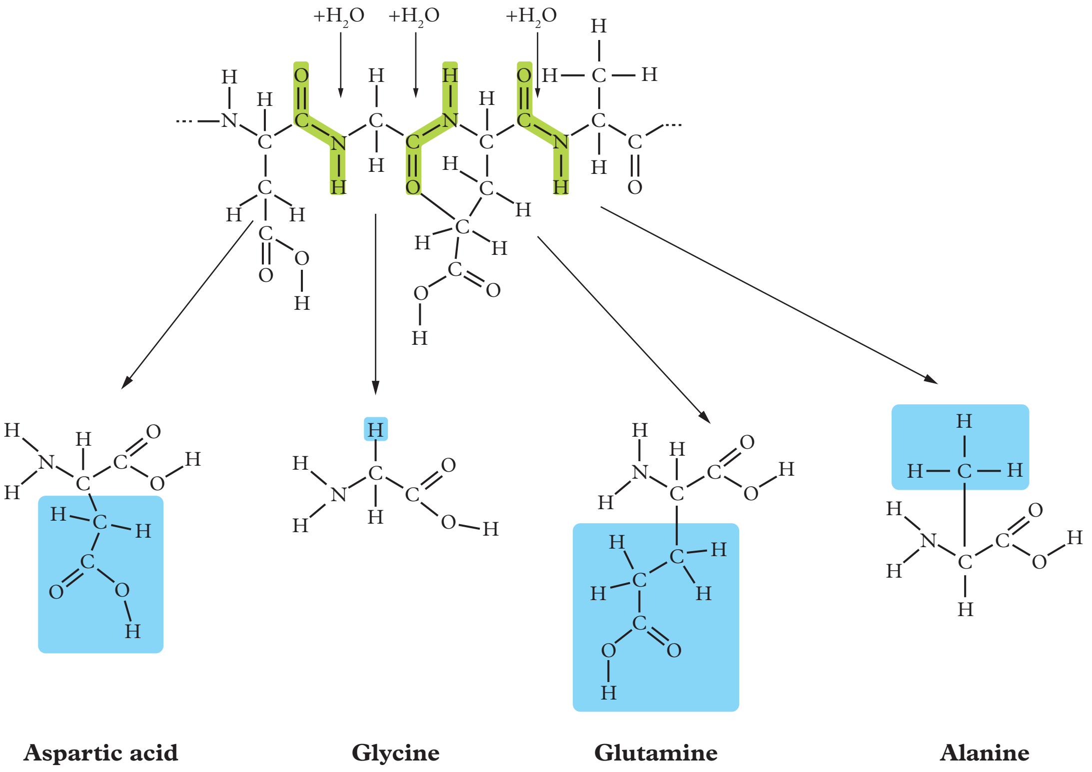


FIGURE 5 Adding water to a protein breaks the peptide bonds and creates amino acids.

Essential amino acids

All 20 amino acids are used to synthesise important metabolic molecules. Humans can naturally synthesise 11 of the 20 α-amino acids. These are considered non-essential amino acids as we do not need to consume them in our diets. The other nine amino acids are the essential amino acids, including histidine, isoleucine, leucine, lysine, methionine, phenylalanine, threonine, tryptophan and valine. They each perform a vital function. For example, leucine supports the production of growth hormones, regulates blood glucose levels and stimulates protein synthesis.

Humans cannot synthesise the essential amino acids; instead, we acquire them from eating protein-rich foods such as beef, poultry, seafood, beans and lentils.

Hydrolysis of proteins

Amino acids can be obtained from proteins via a hydrolysis reaction. The hydrolysis of proteins is when a water molecule is added to each of the peptide bonds holding the protein together (Figure 6). **Enzymes** help catalyse the hydrolysis of protein so that when a water molecule is added to each amide link:

- the covalent bond between the carbon and the nitrogen in the peptide bond breaks
- an –OH from the water is added onto the C=O end of the amino acid to form the carboxyl group
- a H atom is added onto the –NH end to form the amino group of the amino acid.

enzyme
a biological, protein-based catalyst that speeds up chemical reactions

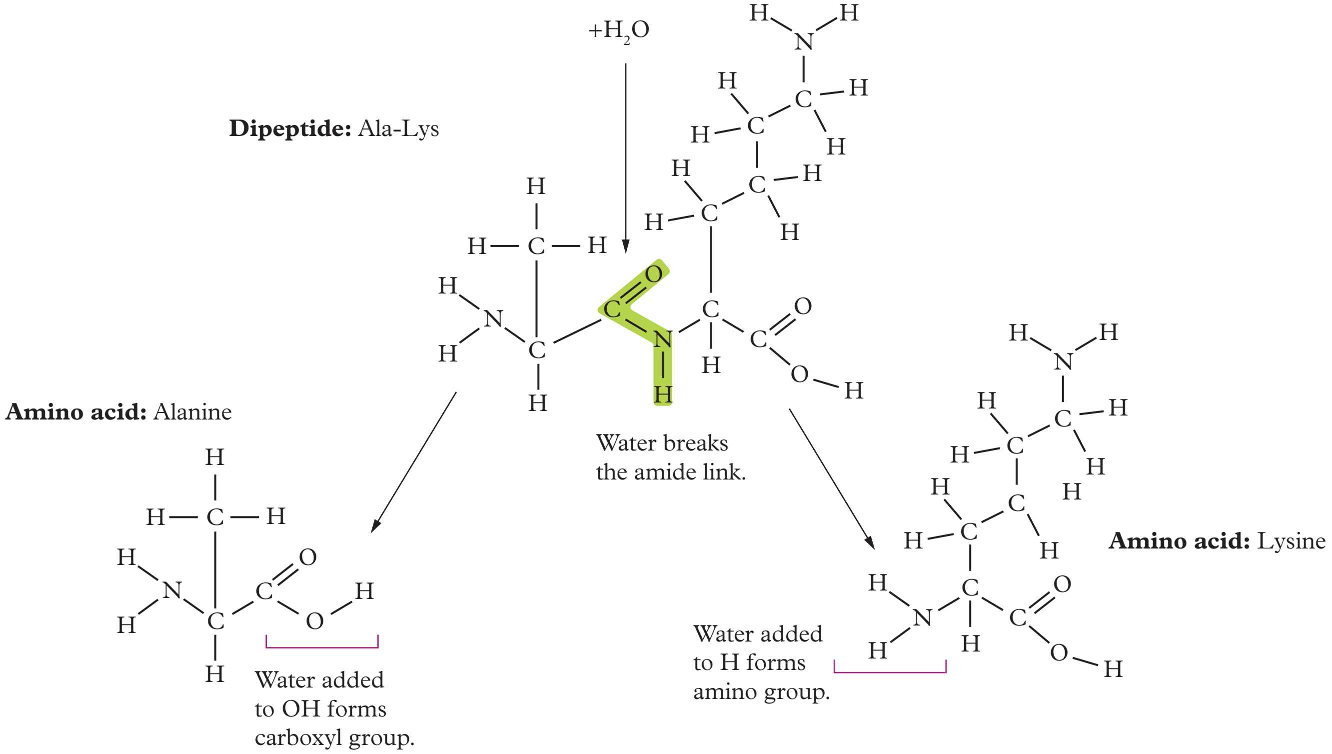


FIGURE 6 Hydrolysis of proteins

Once the small, soluble amino acids have been formed, they can easily move through the bloodstream or gastrointestinal tract into cells to perform different functions. Amino acids can also be synthesised back into proteins.

Condensation reaction to form proteins

In a **condensation reaction**, two molecules combine to form a larger molecule and a small by-product such as water. This is the opposite of what happens in a hydrolysis reaction. An example is shown in Figure 7, where the carboxyl end on one amino acid loses its –OH group, while the amino group of the other amino acid loses a H atom. The –OH and –H form a water molecule, while the carbon of the C=O and the nitrogen of the NH form a covalent bond and become a new peptide bond.

condensation reaction
a reaction in which two molecules with functional groups react to form a larger molecule and a small molecule by-product

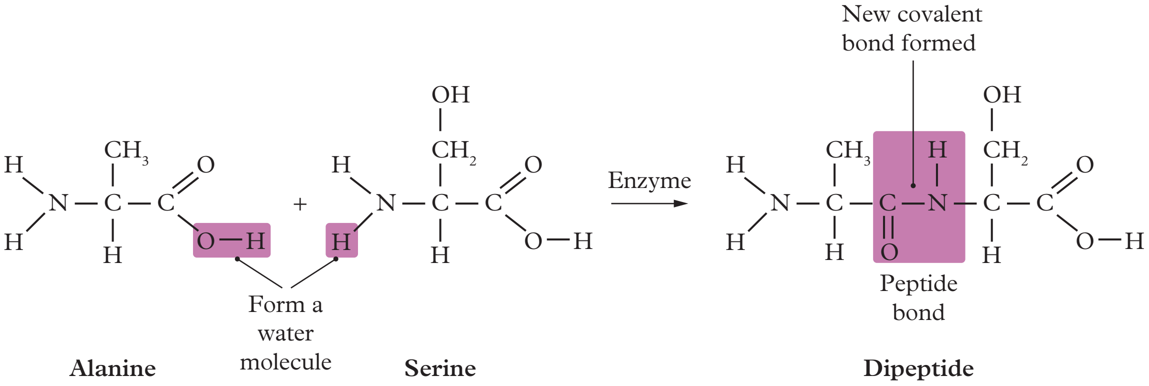


FIGURE 7 Condensation reaction between amino acids

Condensation reactions occurring between the $-NH_2$ group and the –COOH groups on either end of the amino acids allow amino acids to keep linking onto the chain and form the required protein.

FIGURE 8 Summary of the metabolism of proteins in the body

Carbohydrates

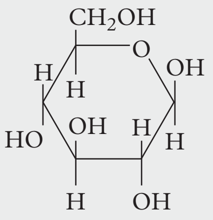
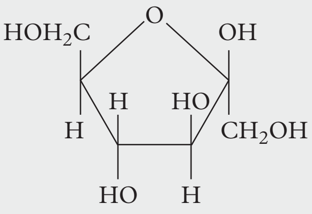
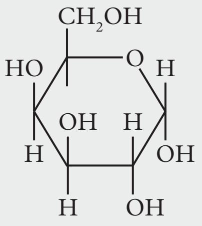
Carbohydrates comprise the elements carbon, hydrogen and oxygen and usually have the general formula $C_x(H_2O)_y$, where *x* and *y* are whole numbers. Carbohydrates are made up of differing sugar molecules called saccharides.

Monosaccharides

Monosaccharides are the building blocks of carbohydrates. They are small, highly soluble and sweet-tasting cyclic molecules. In Table 1 on the next page, you can see three of the most common monosaccharides: glucose, fructose and galactose.

FIGURE 9 Sugars or saccharides are commonly used to sweeten foods.

TABLE 1 Summary of glucose, fructose and galactose

Glucose $C_6H_{12}O_6$	**Fructose** $C_6H_{12}O_6$	**Galactose** $C_6H_{12}O_6$
		
• the most abundant monosaccharide and functions as the key energy source in most forms of life • formed in the cells of green plants by the process of photosynthesis	• the sweetest sugar and is found in fruit juices and honey • not as common as glucose, but it does supply energy to the body in the same way as glucose	• not found as a free monosaccharide; it occurs as a component of larger carbohydrates • obtained by the hydrolysis reaction of the disaccharide lactose (in milk)

You might have noticed that glucose, fructose and galactose molecules are isomers; they have the same molecular formula but different molecular structures. Glucose and galactose are especially similar in their structure, as they both have five chiral carbon atoms. Four carbon atoms are connected to functional groups in the same orientation, but the orientation on one carbon atom differs. You will learn more about chirality in Chapter 12.

ether link
an –O– bond that links together two carbon atoms

glycosidic bond
an ether link between two saccharides

Each carbon ring in a monosaccharide is surrounded by multiple polar –OH functional groups. The –OH functional group can readily form a hydrogen bond with a water molecule, so all small monosaccharides are therefore highly water-soluble. This helps when they are digested by humans, as they can be easily transported throughout the body for use by cells.

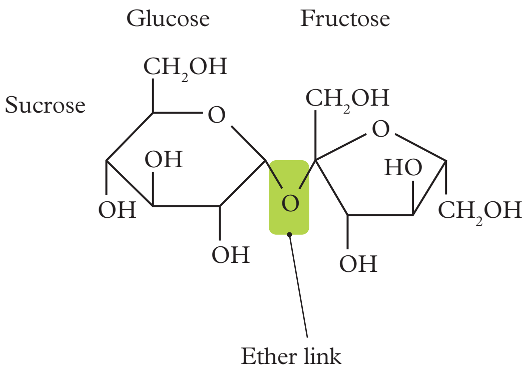


FIGURE 10 The disaccharide sucrose

Disaccharides

Disaccharides are sugars that consist of two monosaccharides joined by an **ether link** (–O–) or **glycosidic bond**. For example, the disaccharide sucrose is made from a glucose and fructose joined by an ether link (Figure 10). Another example is lactose, which is made from glucose and galactose monomers.

Polysaccharides

Polysaccharides are long chains of ether-linked saccharides. For example, the polysaccharides amylose and amylopectin are two different forms of starch; they are both polysaccharides made of numerous linked units of α-glucose monomers (Figure 12).

Amylose is made from long linear chains of α-glucose molecules bonded with an ether link between carbon-1 and carbon-4. All the $-CH_2OH$ groups in amylose have the same orientation. Amylopectin is made from highly branched chains of α-glucose molecules that have additional glycosidic bonds between the –OH group on carbon-1 and carbon-6. The $-CH_2OH$ group is at the top of the α-glucose molecule, off carbon-5.

FIGURE 11 Potatoes are rich in starch.

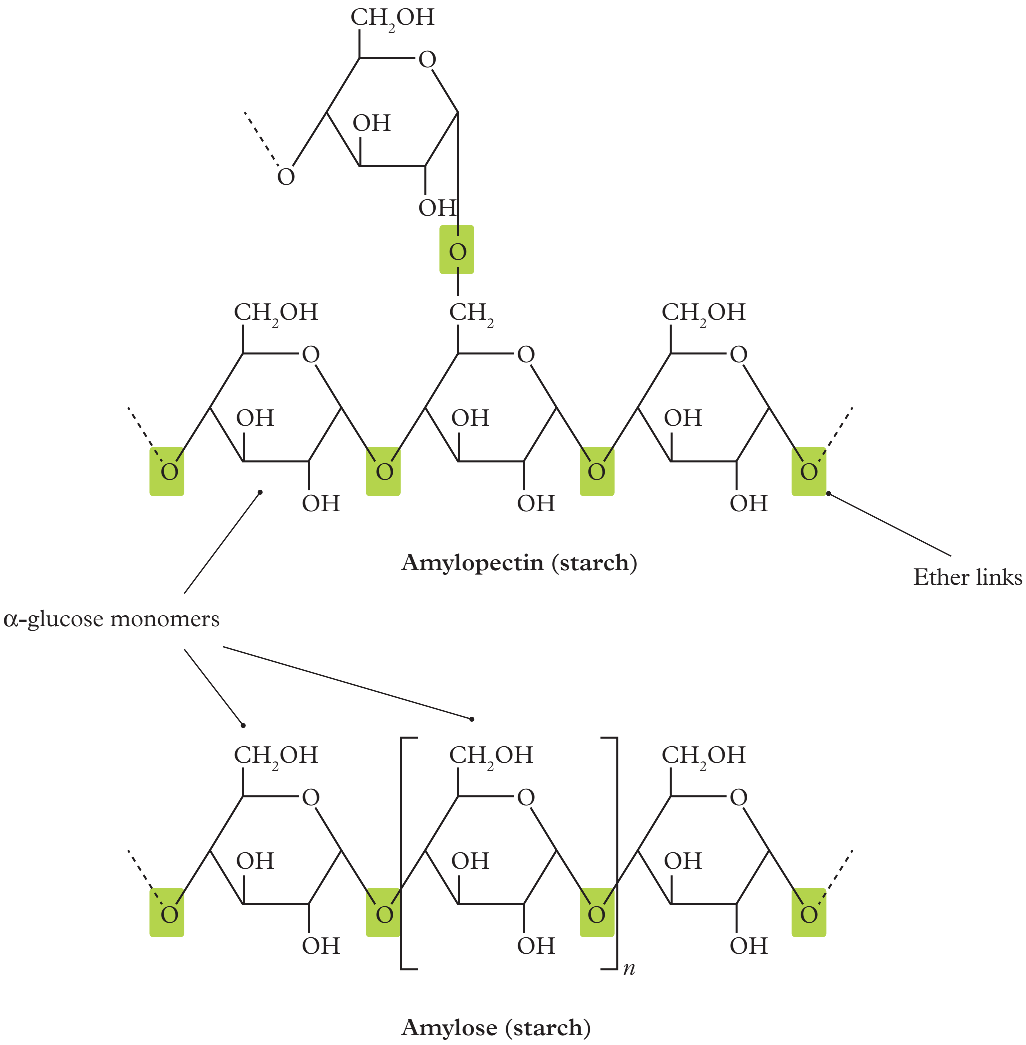


FIGURE 12 Two forms of starch: amylose and amylopectin. Both are polysaccharides made of α-glucose monomers.

If you are confused about the numbering of carbon atoms in a glucose molecule, check out Figure 13. Start numbering the carbon atoms after the ether link: carbon-1 is down from the ether and you travel in a clockwise direction around the molecule

Another common polysaccharide is **cellulose**. Cellulose is a long-chain biopolymer made up of numerous β-glucose units. Figure 14 shows that the orientation of the glucose monomers alternates in cellulose. Humans cannot hydrolyse cellulose as we do not have the specific enzyme, cellulase, to catalyse the reaction. However, we still consume cellulose in our diets, in the form of dietary fibre. Cellulose passes through our systems relatively intact.

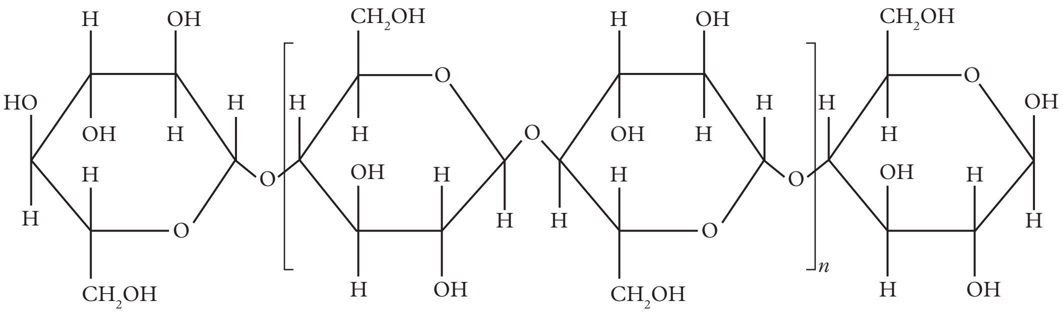
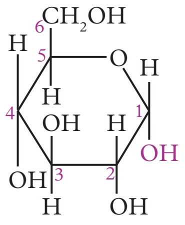
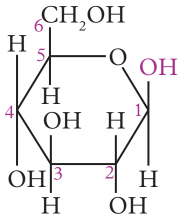


FIGURE 14 The structure of cellulose



Alpha (α) glucose



Beta (β) glucose

FIGURE 13 α-glucose and β-glucose are isomers of glucose that differ in the orientation of OH on the 1-carbon.

cellulose
long-chain linear polysaccharide made up of β-glucose monomers; not digestible by humans

Study tip

An ether link –O– is formed between specific –OH functional groups in all carbohydrates. It is a covalent bond and when it appears between carbohydrate monomers, it is also known as a glycosidic link or bond.

Hydrolysis of carbohydrates

When we consume a source of carbohydrates such as bread, the hydrolysis of starch begins straight away in the mouth. Saliva contains the enzyme amylase, which catalyses the hydrolysis of starch into its smaller monomer, glucose. The catalysed hydrolysis of disaccharides and polysaccharides continues the whole way through the digestive tract.

The hydrolysis of carbohydrates, like all hydrolysis reactions, involves the addition of a water molecule. Figure 15 shows how the addition of a water molecule at an glycosidic bond gives a H atom to one glucose molecule and an –OH to the glucose molecule on the other side of the glycosidic bond. The hydrolysis of the carbohydrate chain continues breaking up the glucose units so they can travel quickly throughout the bloodstream into cells for cellular respiration.

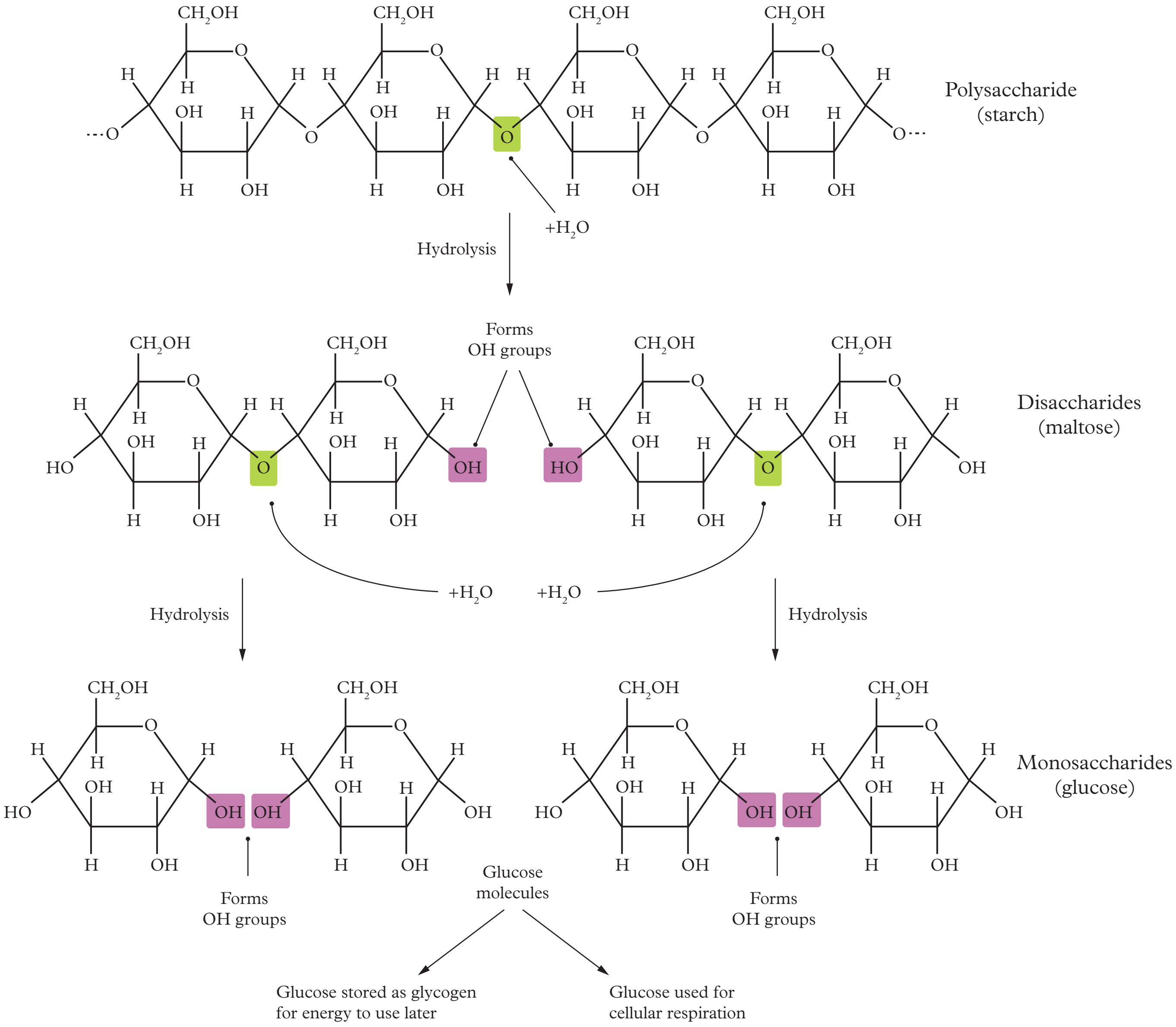


FIGURE 15 Hydrolysis of starch

Most starch-based foods have a combination of amylose and amylopectin. Amylopectin is hydrolysed faster because its branching prevents it from packing as closely as the tightly coiled and linear amylose (Figure 16). Water molecules and amylase can gain access to glycosidic bonds of amylopectin more easily than amylose, so amylopectin digestion occurs faster.

In comparison, amylose is able to closely pack its long linear strands together. This makes it difficult for water and amylase to access and hydrolyse glycosidic bonds.

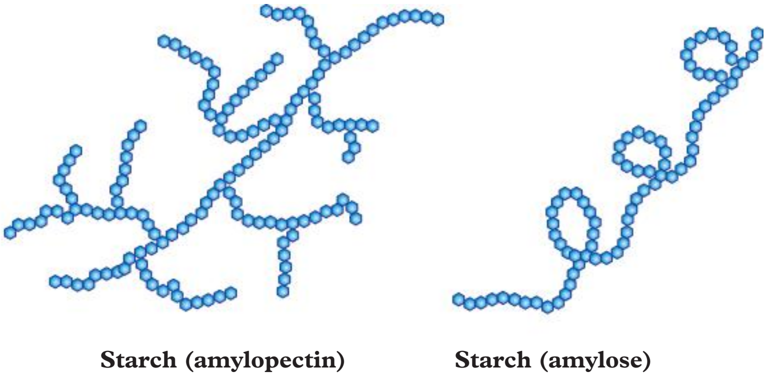


FIGURE 16 The overall structure of amylose and amylopectin

Condensation reactions to form carbohydrates

In Figure 15, you can see that the α-glucose molecules are used in two different ways.

1. Glucose is sent straight to the cells to be converted into energy via cellular respiration.
2. Glucose molecules are stored as glycogen.

Glycogen is the main storage form, and another polysaccharide, of α-glucose. Like amylopectin, it has ether links between carbon-1 and carbon-4, and branching between carbon-1 and carbon-6. However, the glycosidic bond between carbon-1 and carbon-6 occurs twice as often in glycogen as it does in amylopectin. This means glycogen has even more branching than amylopectin (Figure 17).

glycogen
a highly branched polysaccharide made up of glucose monomers; used in animals for energy storage

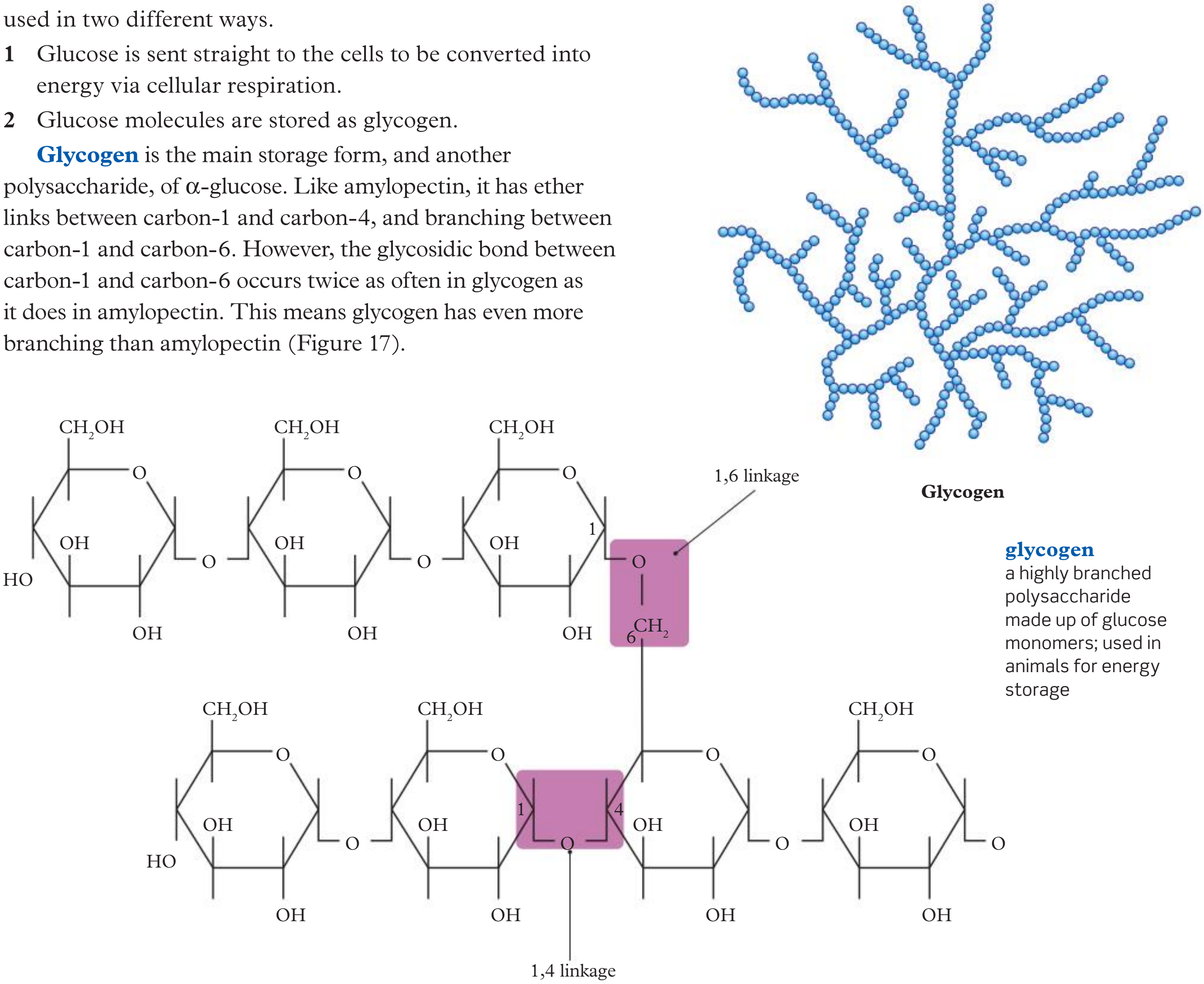


FIGURE 17 Glycogen structure

After large starch polysaccharides have been hydrolysed into α-glucose molecules, the α-glucose molecules that are not used in cellular respiration are stored for later energy use. This occurs via a condensation reaction that converts the glucose monomers back into polysaccharides (in this case, glycogen). The condensation reaction creates a water molecule when two –OH groups from glucose molecules join and form an ether link and a water molecule (Figure 18).

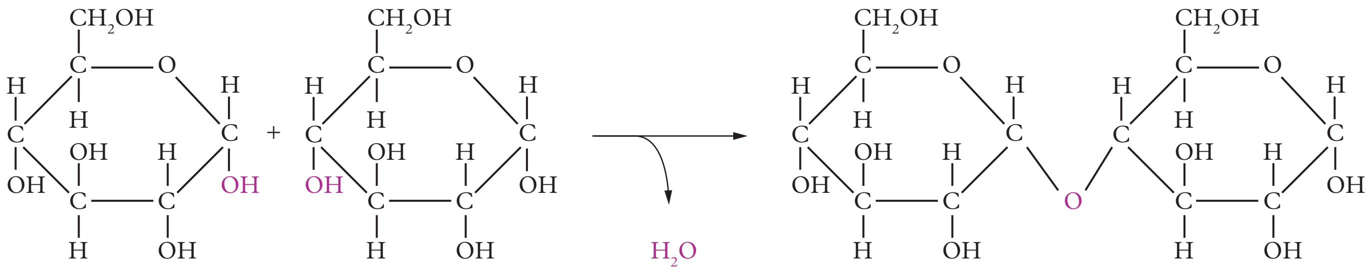


FIGURE 18 Condensation reaction between two α-glucose molecules

The condensation reactions between the α-glucose molecules will continue to occur between the –OH groups on the carbon-1 and carbon-4 atoms and the carbon-1 and carbon-6 atoms to create glycogen. Glycogen can then be hydrolysed again to create glucose for energy use in cellular respiration when required.

FIGURE 19 Summary of the metabolism of carbohydrates in the body

Fats and oils

triglyceride
large non-polar molecule made from three fatty acids and a glycerol molecule

Fats and oils are two common types of lipids also known as **triglycerides**. Triglycerides are made up of carbon, hydrogen and oxygen molecules. They consist of three fatty acid chains (or residues) connected to a glycerol backbone by ester links (Figure 20).

Fats and oils differ in chemical and physical properties. Oils are more common in plants and contain more unsaturated fatty acid residues. Fats are more common among animals and contain more saturated fatty acid residues.

Fats have higher melting points and are solid at room temperature. They can pack together more tightly due to only having C–C single bonds in their chains. Oils have lower melting points because their C=C double bond creates kinks in their chains, which prevents them from packing together as tightly. Since the dispersion forces between the chains with kinks are not as strong, oils are more likely to be liquid at room temperature.

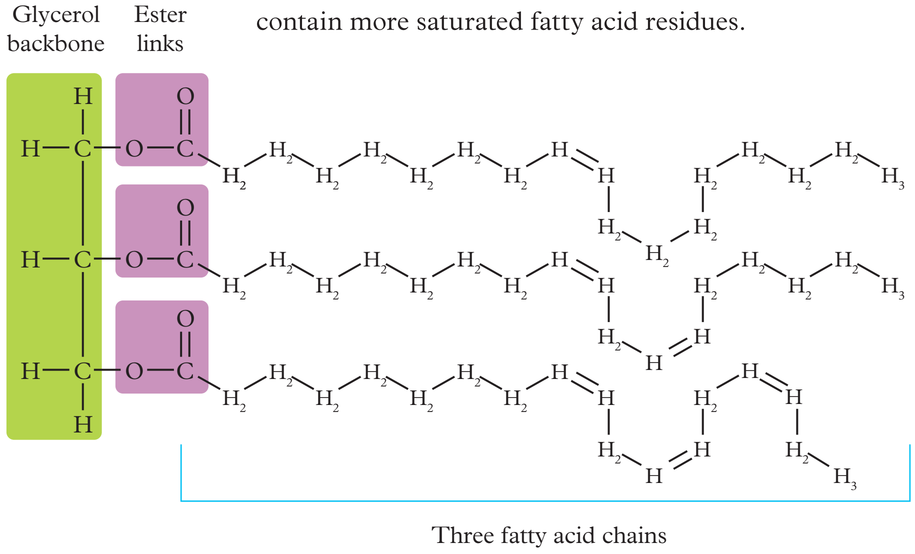


FIGURE 20 Structure of a plant triglyceride containing C=C double bonds

Hydrolysis of fats and oils

Unlike carbohydrates and proteins, lipids are large **hydrophobic** molecules. The digestion process of triglycerides starts in the small intestine. As triglycerides are hydrophobic, the enzyme lipase, which catalyses the hydrolysis reaction, can only react with the surface of a fat globule, making hydrolysis of the triglyceride slow. To increase the rate at which triglyceride hydrolysis occurs, the body uses the digestive fluid, bile. Bile breaks down the large fat globules into smaller molecules, which gives lipase enzymes better access to catalyse the hydrolysis of triglycerides (Figure 21).

hydrophobic
has a low affinity for water; generally non-polar

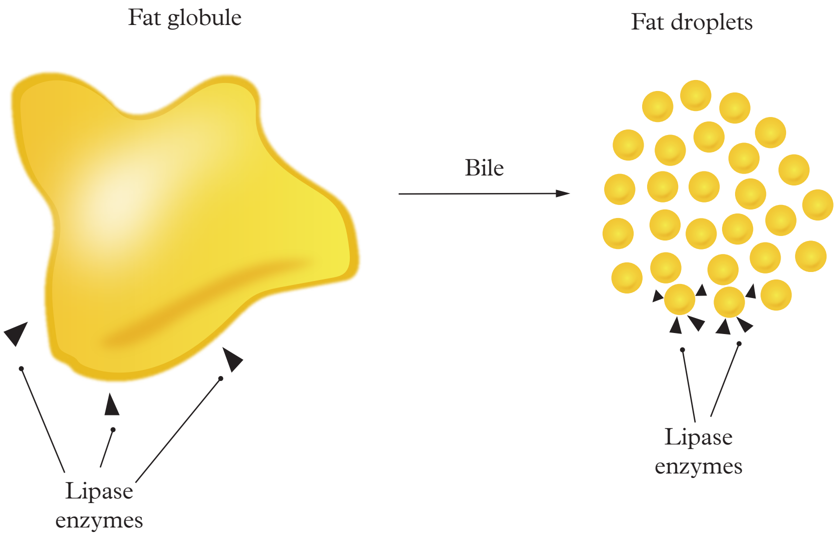


FIGURE 21 Bile breaks down large fat globules into smaller fat droplets so that lipase enzymes can more easily catalyse the hydrolysis of triglycerides.

After bile has broken up and increased the surface area of the fat globule, lipase and water molecules access the triglycerides. The water and enzyme hydrolyse the three ester bonds in the triglyceride to form three fatty acids and a glycerol molecule (Figure 22). Water molecules split and add –H and –OH groups on to form the –COOH of the fatty acid and the –OH of the glycerol. Three water molecules are needed to fully hydrolyse each triglyceride.

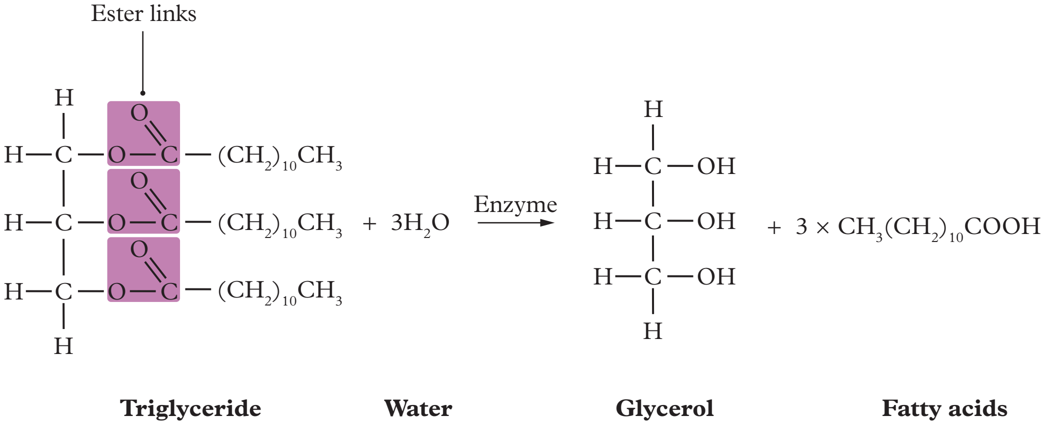


FIGURE 22 Hydrolysis of a triglyceride

After the triglycerides have been hydrolysed into the fatty acids and glycerol molecule, they can be transported through the bloodstream and into the liver, where they can be reformed back into triglycerides via a condensation reaction.

Study tip

When drawing a triglyceride from a fatty acid either from the data book, or in a question, the fatty acid will be listed in the direction $CH_3(CH_2)_{14}COOH$. You must remember that the –COOH group reacts to form the ester link, so start from the end, draw the ester link onto the glycerol backbone and work backwards, drawing the full or semi-structural formula out carefully.

Condensation reactions to form fats and oils

The condensation reaction to reform triclycerides from three fatty acids and a glycerol molecule is the reverse reaction of hydrolysis. The reaction produces three water molecules and a triglyceride that has three ester linkages (Figure 23). The ester links are formed from the reaction between the –COOH group on the fatty acid and the –OH group on the glycerol molecule.

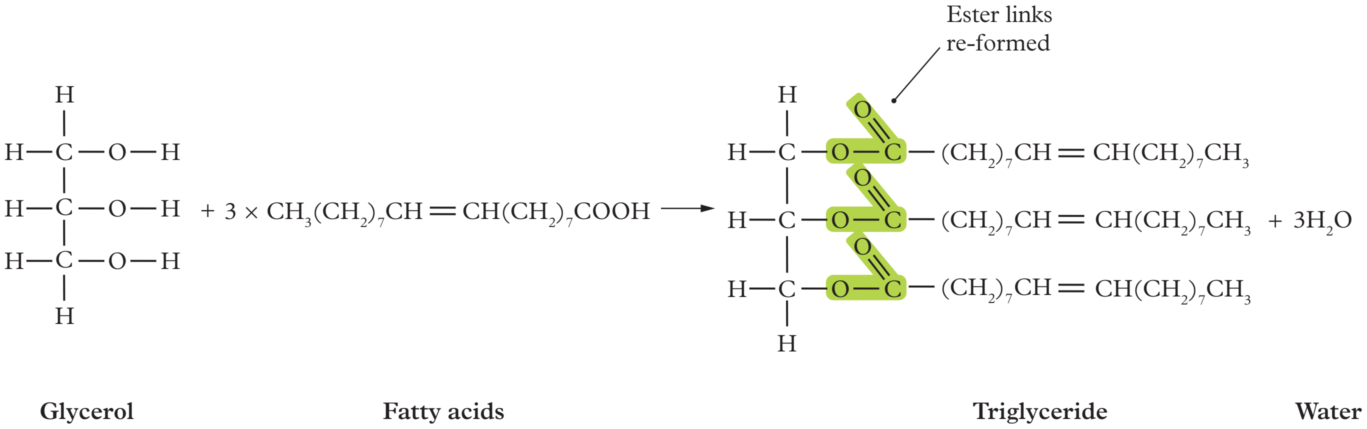


FIGURE 23 The condensation reaction that occurs in the liver to reform a triglyceride from three fatty acids and a glycerol molecule

Once formed, triglycerides are either used as an energy source, or stored in adipose tissue. Adipose tissue is found under the skin or around organs and muscles, and serves as a long-term energy source.

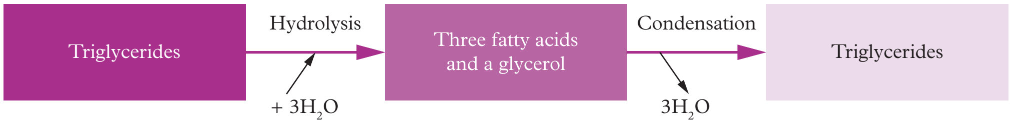


FIGURE 24 Summary of the reactions of lipids

Transesterification of plant triglycerides using alcohols to produce biodiesel

In Chapter 2, you learnt about why biodiesel is a useful and renewable fuel. The same triglycerides that we consume for energy are also used to synthesise biodiesels and used as energy for transport.

Triglycerides can be made up of many different combinations of fatty acids. They can have three of the same or three different fatty acids. Fatty acids in plant oils are more likely to be unsaturated (contain a carbon–carbon double bond). Table 2 shows a selection of common fatty acids and their formulas.

FIGURE 25 Biodiesel can be produced in the lab. The yellow top layer is the biodiesel and the bottom brown layer is glycerol.

TABLE 2 Formulas of common fatty acids

Name	Formula	Semi-structural formula
Lauric	$C_{11}H_{23}COOH$	$CH_3(CH_2)_{10}COOH$
Myristic	$C_{13}H_{27}COOH$	$CH_3(CH_2)_{12}COOH$
Palmitic	$C_{15}H_{31}COOH$	$CH_3(CH_2)_{14}COOH$
Palmitoleic	$C_{17}H_{29}COOH$	$CH_3(CH_2)_4CH_2CH{=}CHCH_2(CH_2)_5CH_2COOH$
Stearic	$C_{17}H_{35}COOH$	$CH_3(CH_2)_{16}COOH$
Oleic	$C_{17}H_{33}COOH$	$CH_3(CH_2)_7CH{=}CH(CH_2)_7COOH$
Linoleic	$C_{17}H_{31}COOH$	$CH_3(CH_2)_4(CH{=}CHCH_2)_2(CH_2)_6COOH$
Linolenic	$C_{17}H_{29}COOH$	$CH_3CH_2(CH{=}CHCH_2)_3(CH_2)_6COOH$
Arachidic	$C_{19}H_{39}COOH$	$CH_3(CH_2)_{17}CH_2COOH$
Arachidonic	$C_{19}H_{31}COOH$	$CH_3(CH_2)_4(CH{=}CHCH_2)_3{=}CH(CH_2)_3COOH$

Study tip

Remember that biodiesels are fatty acid esters.

You may have driven past bright yellow crops of canola during the springtime just before it is harvested. Canola produces an oil that is commonly used to produce biodiesel. Biodiesels are esters – specifically, a fatty acid methyl ester. It is made up from a fatty acid (from the triglyceride) and methanol.

Figure 26 shows a biodiesel molecule made up from oleic acid and methanol.

As each triglyceride has three fatty acids attached to a glycerol backbone, the process of converting a triglyceride into a biodiesel molecule uses the addition of three alcohol molecules. Normally, small alcohols such as methanol are added and then the solution is warmed with a hydroxide catalyst (either potassium hydroxide or sodium hydroxide). This process is known as **transesterification** and is shown in Figure 27. Transesterification produces three biodiesel molecules and one glycerol molecule.

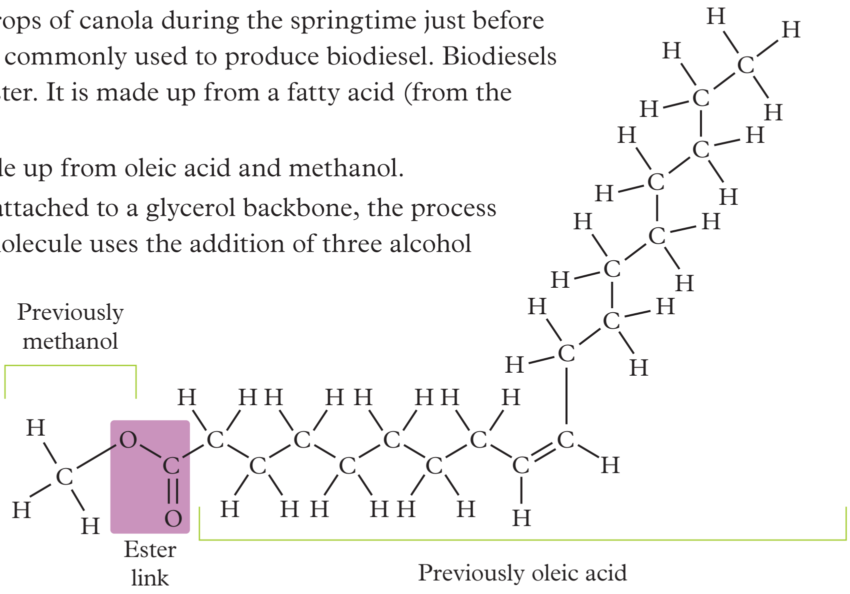


FIGURE 26 A biodiesel molecule that was produced from oleic acid and methanol

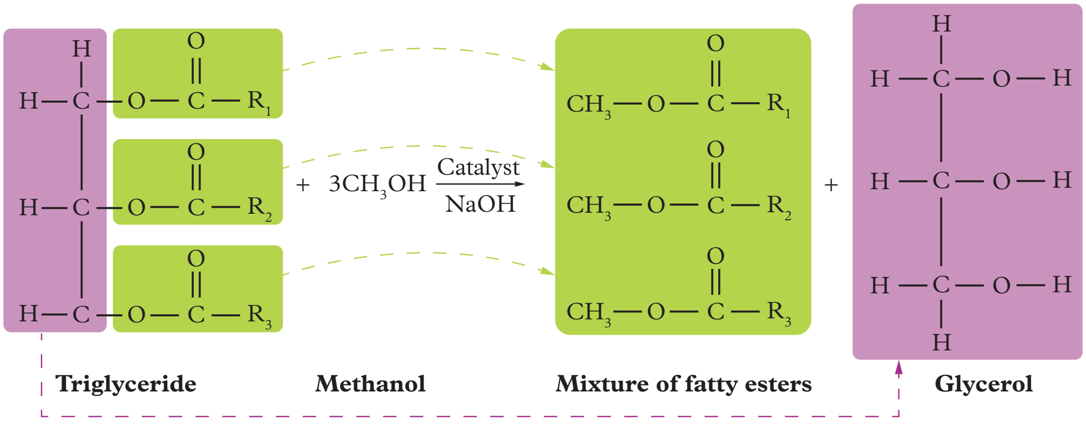


FIGURE 27 General formula of the transesterification process

transesterification the conversion of one fatty acid ester into another fatty acid ester; in this case, a plant triglyceride into a biodiesel molecule

9.2 CHECK YOUR LEARNING

Describe and explain

1 Describe and draw the condensation reaction that would occur between the amino acids Phe, Ala, Asp and Leu.

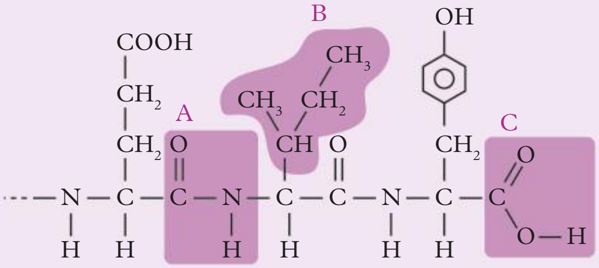
2 Explain the different functional groups that form and draw the products in the condensation reactions between:
  - **a** glycerol and three oleic acid molecules
  - **b** glucose and sucrose
  - **c** methionine, glutamic acid and isoleucine
  - **d** repeating units of β-glucose.

3 Describe and draw the reaction to form biodiesel from the plant triglyceride that contains glycerol and palmitic, oleic and linoleic acids.

FIGURE 28 Production of biodiesel in the laboratory

Apply, analyse and compare

4 Analyse the section of protein shown and answer the following questions.



  - **a** Identify the features labelled A–C in the protein.
  - **b** Identify the three amino acid residues in the protein segment.

5 Consider the structures of cellulose and amylose.
  - **a** Analyse and explain the differences and similarities in the structures.
  - **b** Complete the hydrolysis reactions for both polysaccharides.
  - **c** Explain why humans cannot digest cellulose, but can digest starch.

Design and discuss

6 Discuss the differences and similarities between a hydrolysis reaction and a condensation reaction.

7 Design a flowchart that shows the condensation reactions that form carbohydrates, proteins and triglycerides from their monomers, and the hydrolysis reactions that break the polymers back into their monomers.

9.3

Atom economy and percentage yield

KEY IDEAS

In this topic, you will learn that:

- atom economy is important for chemists to consider when trying to minimise waste
- percentage yield (% yield) can be calculated after you have completed a chemical reaction.

FIGURE 1 Chemists need to design reaction pathways to create their desired product.

Our society relies on the chemical industry to produce a variety of products including medicines, plastics, fuels and foods. Once a chemist determines which product or compound they want, they need to design a chemical reaction pathway that creates the product.

These pathways can be designed to start with widely available and cost-effective materials, such as alkanes and alkenes that can undergo a series of reactions to produce the final product. Chemists must also consider the most efficient or 'green' way to make the desired product. In this topic, we will look at these principles along with simple and complex reaction pathways.

Atom economy

It is important that industry works towards the most efficient industrial processes, with high **atom economy** and low waste. This is important for sustainable development and conservation of natural resources, and is why atom economy is one of the principles of green chemistry.

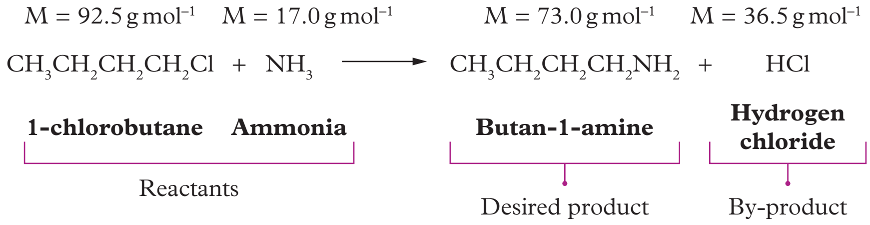
atom economy the ratio of the amount of product yielded to the amount of reactants used in a reaction; a measure of the efficiency of a reaction in converting reactants to products

Atom economy is a measure of how efficient a chemical reaction is. It is based on the ratio of atoms of the desired product to the atoms of the initial reactants. Atom economy is given as a percentage and can be calculated using the equation:

$$\text{percentage atom economy} = \frac{\text{molar mass of desired product}}{\text{molar mass of all reactants}} \times 100$$

The green chemistry principles state that to increase atom economy, reaction pathways should be designed to maximise the incorporation of all reactant materials used in the process into the final product.

Addition reactions have an atom economy of 100% as they only produce one product, the desired product. On the other hand, substitution reactions form a by-product (Figure 2).



$$\underset{\text{M = 92.5 g mol}^{-1}}{CH_3CH_2CH_2CH_2Cl} + \underset{\text{M = 17.0 g mol}^{-1}}{NH_3} \longrightarrow \underset{\text{M = 73.0 g mol}^{-1}}{CH_3CH_2CH_2CH_2NH_2} + \underset{\text{M = 36.5 g mol}^{-1}}{HCl}$$

FIGURE 2 Reaction between 1-chlorobutane and ammonia to form butan-1-amine and HCl

9.3
PRACTICAL: CONTROLLED EXPERIMENT

What is the percentage yield of aspirin? Go to page **516**.

The atom economy of the reaction in Figure 2 can be calculated using the percentage atom equation, where atom economy = $\frac{73.0}{(92.5 + 17.0)} \times 100$. This gives an atom economy of 66.7%.

A high percentage atom economy indicates that the process produces less waste and is therefore more sustainable and aligned to the green chemistry principles. Atom economy can also be calculated for multiple-step reaction pathways. This is done by calculating atom economy one step at a time and disregarding the middle-step product. See Worked example 9.3A for an example of calculating atom economy in a multi-step pathway.

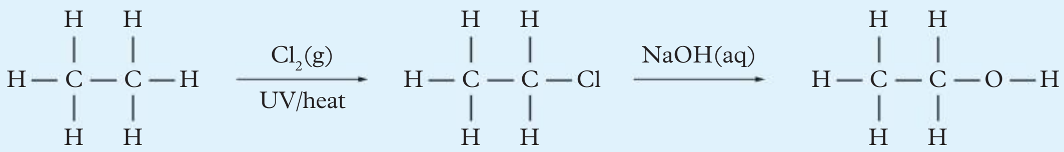
9.3 Challenge
Find me in your obook pro

Then, if you feel confident, try calculating the atom economy of a pathway in Challenge 9.3.

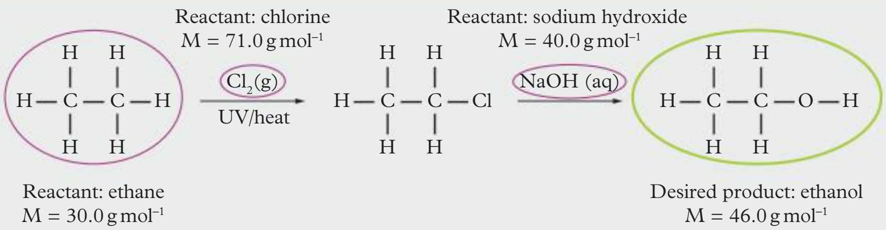
9.3A WORKED EXAMPLE

CALCULATING PERCENTAGE ATOM ECONOMY IN A MULTI-STEP PATHWAY

Calculate the percentage atom economy for the two-step reaction to form ethanol from ethane.



Solution

Think:	Do:
Step 1: Identify your desired product and calculate the molar mass of all the reactants and the desired product.	 Note that chloroethane is not a reactant or a desired product. It is just an intermediary step to the final product, so you do not have to calculate or include this in the calculations. Also note that chlorine and sodium hydroxide are both reactants, so both will need to be calculated and included in the equation.
Step 2: Insert the molar masses into the percentage atom economy equation.	percentage atom economy = $\frac{46.0}{30.0 + 71.0 + 40.0} \times 100$ = 32.6% (3 sig fig)

The importance of high atom economy in industry

Having a high atom economy in reactions is important in industry as it:

- avoids creating waste products and thus prevents pollution
- uses fewer products, which results in lower costs
- uses more sustainable practices
- uses fewer natural resources.

An industrial chemist trying to find a new way to synthesise a product (for example, an ester used in food flavouring) would want to start with the smallest, cheapest reactants they could use. They would then design a pathway that has a minimal number of reaction steps, to keep the atom economy high and not lose reactants at each step. It would also be desirable to produce the lowest number of waste products at each step, as getting rid of waste is costly and can harm the environment.

Percentage yield

Most chemical reactions completed in industrial facilities are not 100% efficient. **Percentage yield** is calculated based on the completion of a reaction. It requires you to know the **actual yield** (or mass) of a final reaction, and the **theoretical yield** that is the ideal outcome for the reaction. Theoretical yield can be calculated using **stoichiometry**.

The equation for percentage yield is:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

percentage yield comparison of actual yield and theoretical yield from a completed reaction

actual yield the final measured mass of product after completing a reaction

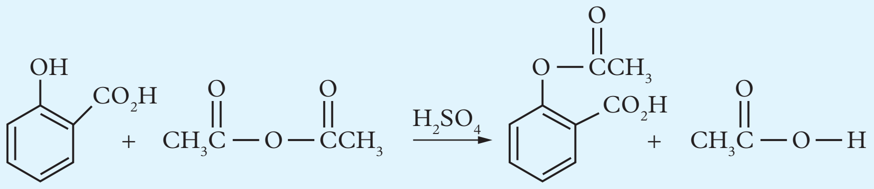
theoretical yield the amount of products expected to be formed from a reaction, calculated using stoichiometry

stoichiometry calculations using mole ratios for a chemical reaction

9.3B WORKED EXAMPLE

CALCULATING PERCENTAGE YIELD

Students completed a reaction to make aspirin. They reacted 3.00 g of salicylic acid with an excess amount of acetic anhydride. This produced 2.71 g of aspirin after the reaction.



Salicylic acid **Acetic anhydride** **Aspirin** **Ethanoic acid**

a Calculate the theoretical yield of aspirin from the reaction.

b Calculate the percentage yield of aspirin the students produced in their reaction.

Solution

Think	Do
Step 1: Calculate the theoretical yield using stoichiometry. We know that salicylic acid is the limiting reactant, as the question says there is an excess of acetic anhydride, so we do not have to find the limiting reactant here. M(salicylic acid) = 138.0 g mol^{-1} M(aspirin) = 180.0 g mol^{-1}	**a** $n(\text{salicylic acid}) = \frac{m}{M}$ $= \frac{3.00}{138.0}$ $= 0.0217\text{ mol}$ $n(\text{aspirin}) = n(\text{salicylic acid})$ $= 0.0217\text{ mol}$ $m(\text{aspirin}) = n \times M$ $= 0.0217 \times 180.0$ $= 3.91\text{ g}$ (3 sig fig)
Step 2: Calculate the percentage yield of aspirin that the students produced in their reaction.	**b** $\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$ $= \frac{2.71}{3.91} \times 100$ $= 69.3\,\%$ (3 sig fig)

Study tip

When calculating % yield, don't forget to check if there is a limiting reactant in the reaction. The limiting reactant is the one with the lowest amount (*n*), and it will determine the maximum amount of product possible. The mass of the final product is calculated using the limiting reagent.

Percentage yield for a pathway

You can also calculate overall percentage yield for a multi-step reaction pathway. As you move through each reaction in a pathway, the percentage yield of each reaction will have an overall effect on the final percentage yield. This means that you must calculate the percentage yield for each step and then multiply them all together to obtain the final percentage yield.

$$\text{percentage yield} = \frac{\text{actual yield of reaction 1}}{\text{theorectical yield of reaction 1}} \times \frac{\text{actual yield of reaction 2}}{\text{theorectical yield of reaction 2}} \times 100$$

If the percentage yield of each step is given, you can just multiply each of those together.

$$\text{overall percentage yield} = \frac{\text{percentage yield of reaction 1}}{100} \times \frac{\text{percentage yield of reaction 2}}{100} \times 100$$

High percentage yields are important in industry as they limit waste and increase production. The factors that affect the percentage yield in a reaction or a pathway are:

- when not all reactants are completely reacted
- if some reactants/products are ‘lost’ during the reaction
- when equilibrium reactions do not proceed to completion
- if the rate of reaction is too low
- if there are multiple inefficiencies in pathways.

9.3 CHECK YOUR LEARNING

Describe and explain

1 Explain the difference between percentage atom economy and percentage yield.

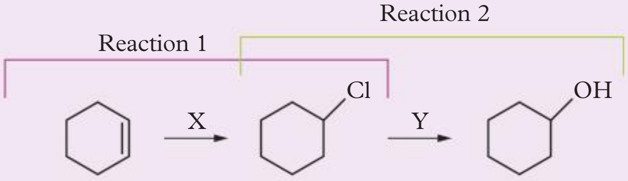
2 Calculate the overall percentage yield for the reaction pathway below.

Compound X → Compound Y → Compound Z

Reaction X → Y yields 83% and reaction Y → Z yields 47%.

Apply, analyse and compare

3 Analyse the pathway shown and answer the following questions.



Cyclohexene **Chlorocyclohexane** **Cyclohexanol**

a Identify each of the reactions that are occurring.

b Identify the reactants, catalyst and conditions at points X and Y.

c Compare the atom economy for reaction 1 and reaction 2.

4 Calculate and compare the percentage atom economy for the two reactions to form propanol below.

$$CH_3CH_2CH_2Cl + NaOH \rightarrow CH_3CH_2CH_2OH + NaCl$$

$$CH_3CH_2CH_2Br + NaOH \rightarrow CH_3CH_2CH_2OH + NaBr$$

Design and discuss

5 From your answers in Question 4, discuss which reaction better adheres to green chemistry principles.

6 Design a pathway with the highest atom economy to form the ester ethyl hexanoate with starting reactants ethene and hexane.

7 Design a flowchart of all the possible pathways and reactions that ethene could complete. Calculate the atom economy for each and comment on which would best adhere to the green chemistry principles.

8 Discuss whether it is more important to have a high percentage yield or a high atom economy. Justify your reasoning.

9.4 Sustainability of the production of chemicals

KEY IDEAS

In this topic, you will learn that:

- green chemistry principles are designed to help chemists work toward Sustainable Development Goals
- renewable feedstocks are used to move away from our reliance on fossil fuels
- catalyst development and selection and atom economy work together to minimise waste products
- chemical design should aim to increase the efficiency of reactions and decrease the toxicity of products produced.

To help us achieve Sustainable Development Goal 12: Responsible consumption and production, we must consider three principles of green chemistry:

- **Use of renewable feedstocks:** Raw materials or feedstocks should be made from renewable (mainly plant-based) materials, rather than from fossil fuels whenever practicable.
- **Catalysis:** Catalysts should be selected to generate the same desired product(s) with less waste and using less energy and fewer reagents in reaction processes/pathways.
- **Designing safer chemicals:** Chemical products should be designed to achieve their intended function while minimising toxicity.

Use of renewable feedstocks

Fossil fuels are used to produce many of our organic chemicals, fuels, energy sources and consumer products. They are not renewable, and we are currently using them at an unsustainable rate. As the global population rises and the demand for products made from fossil fuels increases, the switch to renewable feedstocks is essential.

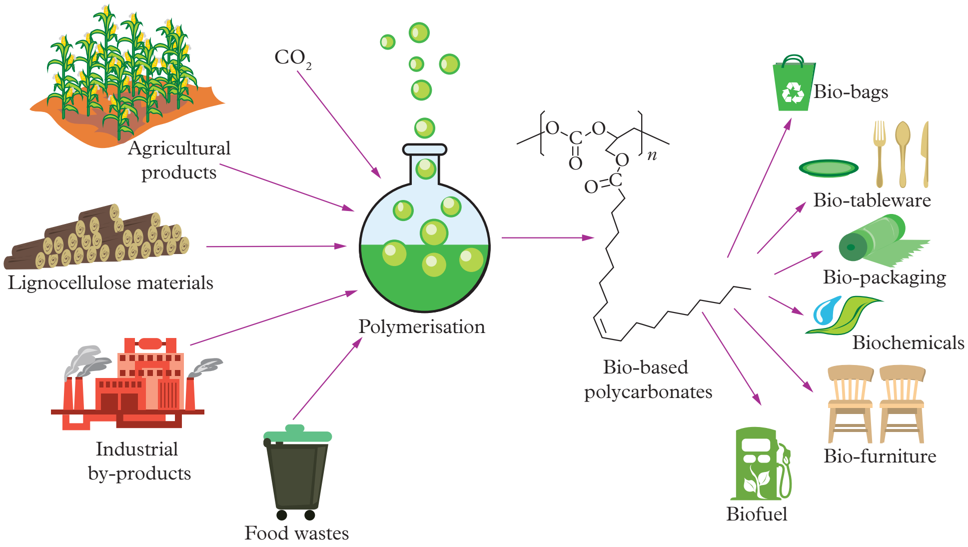


FIGURE 1 Use of renewable feedstocks to create bio-products

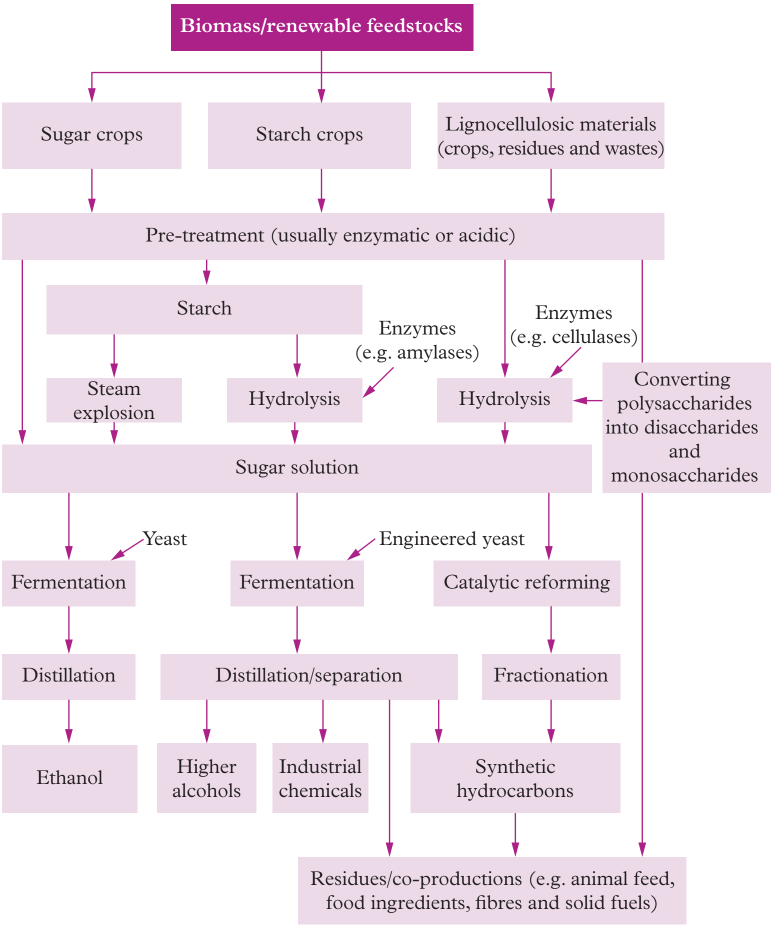


FIGURE 2 Flowchart for the production of biochemicals and biofuels from biomass and renewable feedstocks

Carbon from the air is removed as CO_2 and used by many plants, crops and algae. These plants can be harvested and used as biomass. Biomass can be the feedstock (raw material) used as a starting point for new development of chemical products, depending on what is needed.

The key to developing new processes and chemical pathways using renewable feedstocks is to make sure they:

- are carbon neutral; the amount of carbon put in is equal or greater than the amount of carbon put out
- require low energy input
- are non-toxic to humans and the biosphere
- produce minimal waste.

Significant advances have already been made in the development of fuels, materials and chemicals from renewable feedstocks. Many of these have already made their way into consumer products and are currently being used. The general process for converting the renewable feedstock biomass to bio-hydrocarbons can be seen in in Figure 2.

Fuels from renewable feedstocks

Biodiesel produced through the transesterification of plant oils is a renewable feedstock. This is one product that we are adapting to use in everyday society, over the use of fossil fuel-derived diesel.

There are also biofuels, such as bioethanol and biobutanol from sugars, that originate from biomass such as corn, cereal and sugar beet crops, and algae.

Bioethanol and biobutanol are both produced from biomass by hydrolysis and sugar fermentation. Biomass contains a complex mixture of carbohydrate polymers such as cellulose, hemicellulose and lignin.

The first step in the production of bioethanol and biobutanol is to break down the biomass and create sugars. This is done by treating the biomass with enzymes or acids to reduce the size of the feedstock and open up plant structures.

The cellulose and hemicellulose are then hydrolysed into sucrose (a disaccharide) and fermented into bioethanol. Lignin is often used as a fuel for the ethanol plant boilers and not converted into bioethanol.

Three processes are used to convert biomass into sugars used for fermentation into bioethanol. These three processes are: concentrated acid hydrolysis, dilute acid hydrolysis and enzymatic hydrolysis (Figure 3). Each process has benefits and drawbacks depending on how large the process is and what the final product is used for.

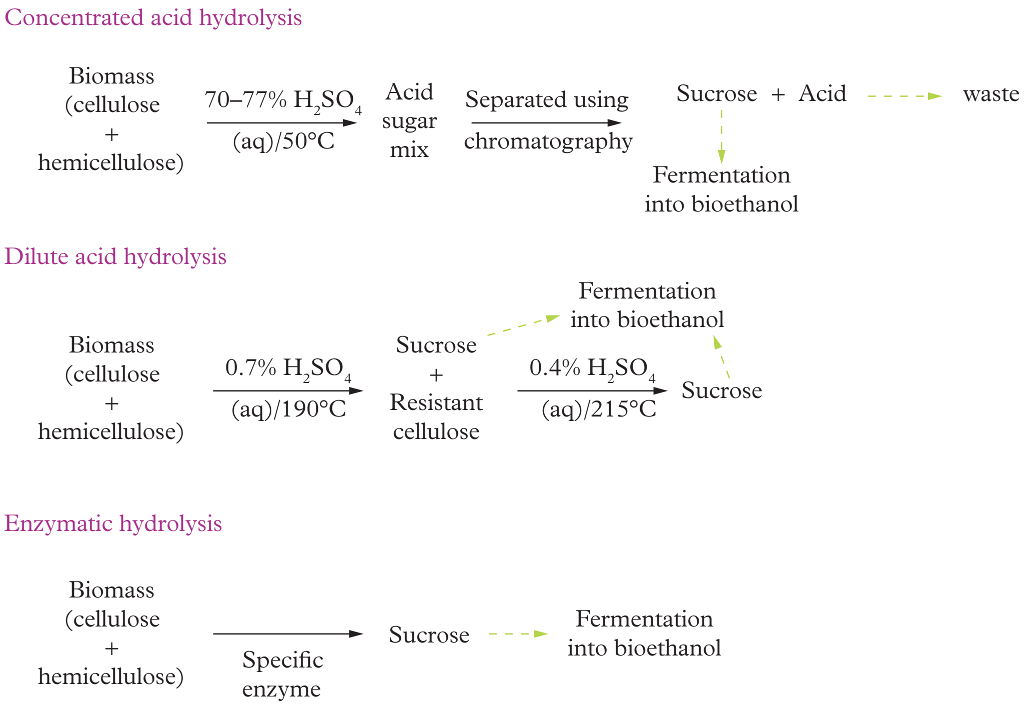


FIGURE 3 The process of converting biomass into sucrose for fermentation into bioethanol

Once the breakdown from the biomass into sucrose is completed, it is then fermented into ethanol. Yeast is added; it contains the catalytic enzymes to break sucrose into fructose and glucose and to ferment the monosaccharides into ethanol (Figure 4).

$$\text{Sucrose} \xrightarrow[\text{catalyst}]{\text{Invertase}} \text{Glucose} + \text{Fructose}$$

$$\text{Glucose or fructose} \xrightarrow[\text{catalyst}]{\text{Zymase}} \text{Two ethanol molecules} + \text{Two } CO_2 \text{ molecules}$$

FIGURE 4 The sugar fermentation process

The fermentation process takes around three days to complete and produces CO_2 as a by-product. Once fermentation is complete, the ethanol mixture contains a significant amount of water that needs to be removed. This is achieved by using fractional distillation, which you will learn more about in Chapter 10. The ethanol mixture is heated and, as ethanol has a lower boiling point (78.3°C) than water (100°C), ethanol turns into vapour and is separated before water can be condensed.

Biobutanol production uses the anaerobic microbial species *Clostridium acetobutylicum* to produce a mixture of acetone-butanol-ethanol (ABE) from biomass. ABE has a ratio of three parts acetone, six parts butanol and one part ethanol. It uses the same general process as that used for converting biomass into bioethanol; however, it uses a two-part fermentation process involving an acid fermentation phase and a solvent fermentation phase. Algae has also been trialled as the biomass for this process. Figure 5 on the next page shows the three different biomass feedstocks that have been used to produce biobutanol by ABE fermentation.

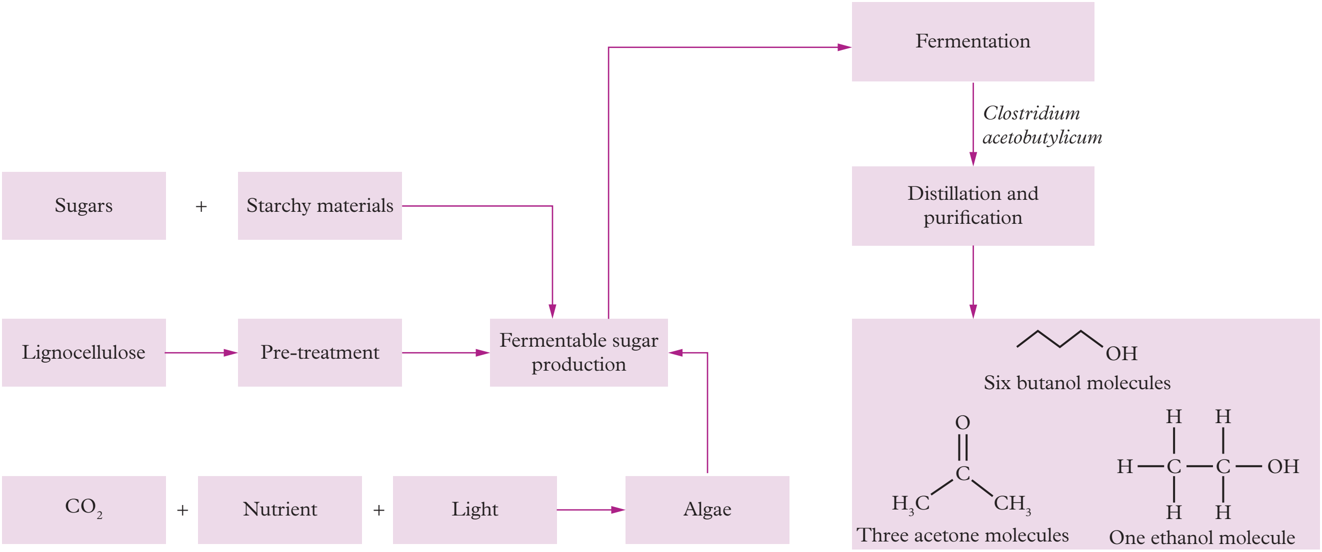


FIGURE 5 The ABE fermentation process using three different feedstocks

TABLE 1 Energy content of common biofuels

Fuel	Energy content
Petrol	32 MJ/L
Biobutanol	29.2 MJ/L
Bioethanol	19.6 MJ/L

Bioethanol and biobutanol can both be used as an addition to petrol for transport. However, the downside is that these fuels have a lower energy content than octane, which is currently the main constituent in petrol. You can see the energy contents of petrol, biobutanol and bioethanol in MJ/L in Table 1.

The chain length and hydroxyl group in butanol give it a similar viscosity to normal petrol. Different percentage blends of petrol, ethanol and butanol are currently being tested to enable cars to run on more renewable fuels.

Plastics from renewable feedstocks

bio-based polymers
biomass-derived polymers or plastics

biodegradable plastics
polymers or plastics that can be broken down by living organisms

Bio-based polymers are plastics made from biomass, but do not confuse them with **biodegradable plastics**. Unlike bio-based polymers, biodegradable plastics do not always come from renewable sources. Biodegradability refers to chemical structure and the ability to break down over time. A bio-based polymer is not necessarily biodegradable – that will depend on its chemical structure. You can see the crossover between the two terms in Figure 6.

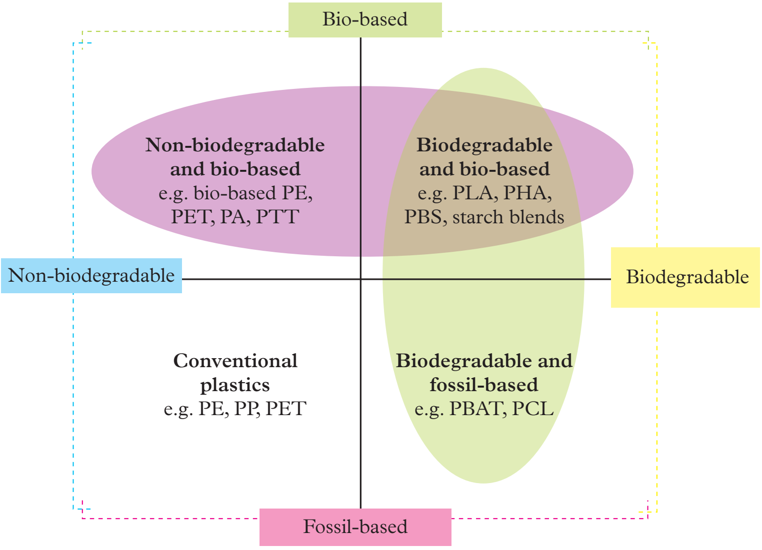


FIGURE 6 The difference between biodegradable and bio-based polymers

Study tip

IUPAC suggests that the use of the term *bioplastic* should be discouraged. Use the expression *bio-based polymer* instead of bioplastic.

There are many bio-based polymers, and more are being created each day as companies begin to understand the benefits of using a plastic made from a renewable source. The process of making bio-based polymers is not that different from the process of making polymers from fossil fuels. The two processes shown in Figure 7 only differ by one step. Production of biopolyethylene requires the conversion of ethanol from biomass into ethene; steam cracking does not have this step, as it produces ethene straight from fossil fuels.

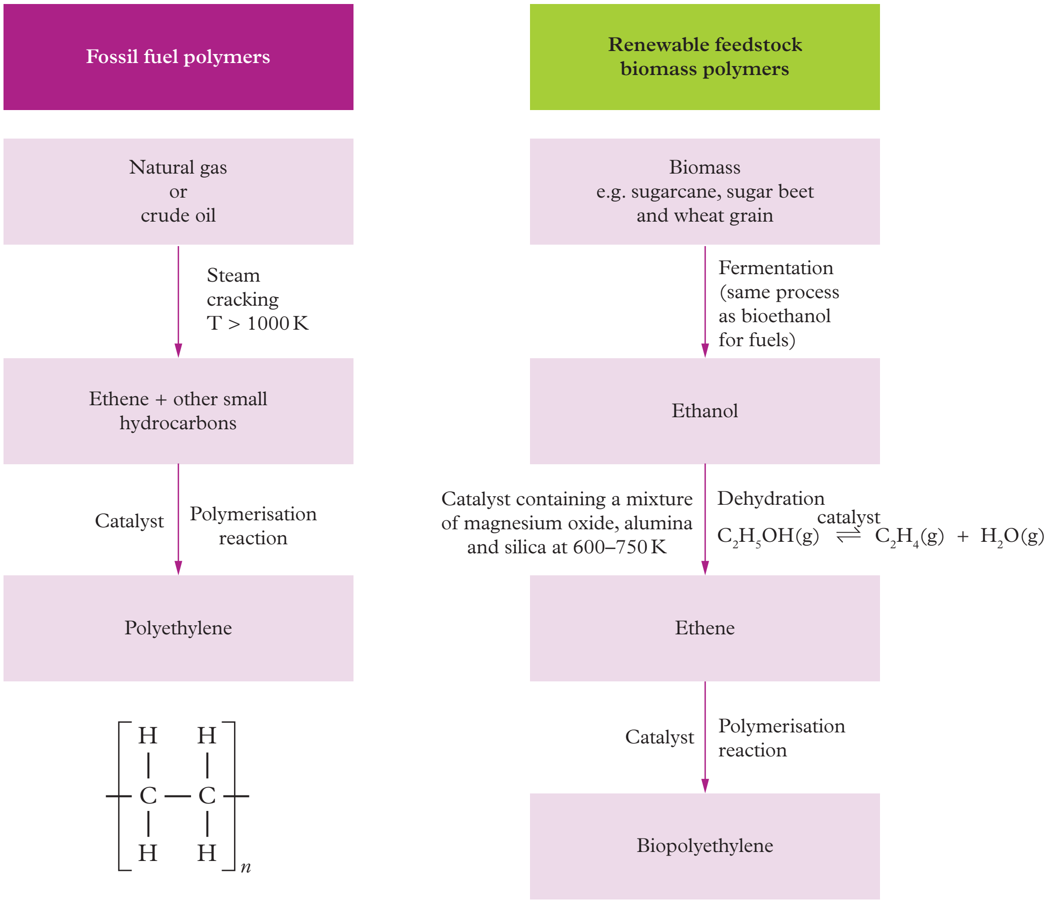


FIGURE 7 The processes for production of polyethylene from fossil fuels and biomass

The main benefit of using biomass to produce polyethylene is that growing crops, specifically sugar cane, separates around 2.15 tonnes of CO_2 for photosynthesis per tonne of biopolyethylene produced. Biopolyethylene, like polyethylene, is not biodegradable, but the same recycling streams can be used for both. The main disadvantage of producing biopolyethylene is that intensive agricultural practices are required to produce the crops, including the use of fossil fuels for transport, and fertilisers and pesticides for growing the crops.

FIGURE 8 Sugar cane is commonly used as a renewable feedstock to produce fuel.

Catalysis

catalysis
the acceleration of a chemical reaction by a catalyst

Study tip

Incorporating catalysts into a reaction pathway can also address the green chemistry principles of *Design for energy efficiency* and *Atom economy*. Can you think of any other links between the principles?

Catalysis is the acceleration of a chemical reaction by a catalyst. Catalysts provide an alternate reaction pathway with a lower activation energy and are not consumed in the process. This means that they can be used in small amounts and, in principle, be reused and recycled indefinitely.

For years, percentage yield has been the main driving factor for chemical processes, leading to the disregard of chemical waste produced or reactants required. This needs to shift to focus on reducing waste and selecting reactants and catalysts for reactions that minimise waste. Atom economy also has its part to play here because, with the selection of a good catalyst, reactants can be chosen so that atom economy is higher and waste products are minimal.

For example, let's look at the reduction reaction of a ketone into a secondary alcohol, in Figure 9. The original reaction has an 81% atom economy, meaning that it has waste products.

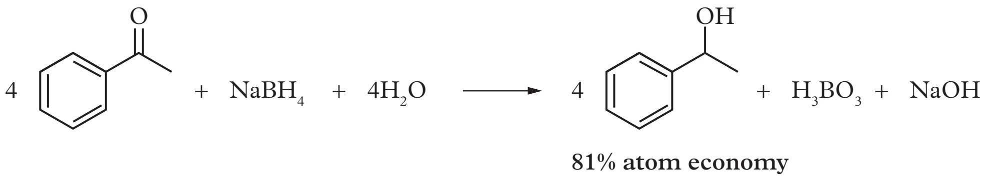


FIGURE 9 Reduction of a ketone to a secondary alcohol, with 81% atom economy

So, how can this reaction improve according to the principles of green chemistry? An atom economy of 100% would mean that there are no waste products. Reducing a ketone to a secondary alcohol requires the addition of H_2. However, under normal conditions, hydrogen will not react with a ketone to any extent. The solution here is to involve a catalyst in the reaction – in this case, palladium-on-charcoal (Pd-on-C) is the catalyst of choice. With the catalyst now involved in the reaction, you can see the hydrogenation of the ketone in Figure 10 with 100% atom economy.

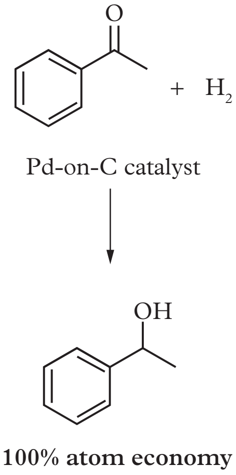


FIGURE 10 Reduction of the same ketone to the same secondary alcohol with the use of a catalyst, giving 100% atom economy

Another benefit of catalytic hydrogenation, beside the reduction of waste products, is that hydrogen is an inexpensive reactant. This process is used commonly in the petrochemical industry.

Pharmaceutical and chemical industries are moving towards the use of catalysts to increase the rate of reactions and minimise the production of waste products. Current research into the different uses of catalysts is exploring how heterogeneous, homogeneous organic catalysts and enzymes can be used. Enzymes are specific to one biochemical reaction and operate in mild conditions, but researchers are finding ways to use their unique applications in pharmaceutical and other chemical industries.

TABLE 2 Advantages and disadvantages of enzyme-catalysed production of biodiesel

Advantages	Disadvantages
• Use of low-quality lipid feedstocks • Low energy input • Generation of pure quality biodiesel and glycerol • Easy recovery of products • Less water wasted	• High cost • Long reaction time • Inhibition of enzymes due to alcohol and glycerol content

For example, biodiesel production through the transesterification of plant triglycerides will often use an inorganic hydroxide catalyst. Researchers are now looking at using a lipase enzyme catalyst instead. As with all chemical processes, there are advantages and disadvantages to the enzyme-catalysed production of biodiesel from plant triglycerides over the current inorganic catalyst method. These advantages and disadvantages are presented in Table 2.

If research can overcome these disadvantages, then the enzyme-catalysed production of biodiesel will be greener and a more sustainable approach to large-scale biodiesel production for future green energy needs.

Designing safer chemicals

The green chemistry principle of designing safer chemicals aims to ensure that chemicals designed are safe and efficient. The definition of toxicity of a chemical has two parts: toxicity to humans (adverse side effects and unintended reactions) and toxicity to the environment (reactions with unintended biological targets). This means that chemists not only have to understand chemistry, but also toxicology and environmental science so that they can design safer chemicals, either to replace currently used toxic chemicals or create brand new chemicals.

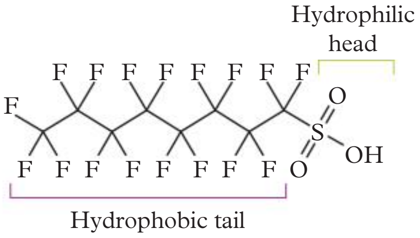


FIGURE 11 Example of a fluorosurfactant

Designing safer chemicals for firefighting foam

Firefighters use firefighting foam to help suppress fires. Some of the current foams used contain fluorosurfactants. You can see an example of a fluorosurfactant in Figure 11. There are many different types of fluorosurfactants, but generally, they are long carbon-based chains surrounded by fluorine atoms.

The downside to fluorosurfactants is that they are bioaccumulative – they build up and persist in the environment.

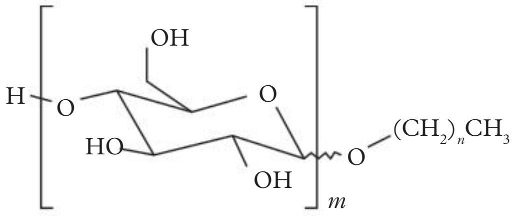


FIGURE 12 An example of an alkyl polyglycoside, where *m* and *n* are whole numbers

One company has started looking at using organic surfactants that are already used in household items (such as soaps and toothpaste), such as alkyl polyglycosides and alkyl sulfates.

These organic surfactants are derived from plant sugars and oils, rather than fossil fuels, which makes them adhere to the green principle of using renewable feedstocks. They are also non-toxic, are not environmentally persistent or bioaccumulative. The newly developed firefighting foam also contains complex carbohydrates and sugars, among other additives. Tests have shown that it is biodegradable, completely falling apart after six weeks.

Many household cleaners are heading in the same direction. The companies using combinations of all the green chemistry principles will eventually help us in achieve Sustainable Development Goal 12: Responsible consumption and production.

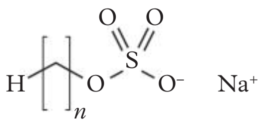


FIGURE 13 An example of the structure of an alkyl sulfate, where *n* is a whole number

9.4 CHECK YOUR LEARNING

Describe and explain

1 Describe how atom economy relates to catalyst selection.

2 Explain what a renewable feedstock is and how it is different to fossil fuel as a feedstock.

3 Explain the disadvantages of using bioethanol and biobutanol in fuels for transport.

Apply, analyse and compare

4 Compare the terms *biodegradable* and *bio-based polymer*.

5 Compare the production of bio-based polymers with the production of a conventional fossil fuel–based polymer, including the advantages and disadvantages of each process.

Design and discuss

6 Define the three green chemistry principles explored in this chapter and discuss how each links to Sustainable Development Goal 12.

7 Design a flowchart showing the steps for the synthesis of a fuel from biomass.

8 Research why IUPAC has suggested the use of the term *bio-based polymer* instead of *bioplastic*.

CHAPTER 9 Review

Chapter summary

9.1

- Organic compounds can form new compounds through substitution, addition, oxidation, condensation and hydrolysis reactions.
- Some reactions require catalysts and specific conditions to occur.
- An organic reaction pathway is created to synthesise a more complex molecule from smaller cheaper starting molecules; these pathways will have multiple steps.

9.2

- Carbohydrates, proteins, fats and oils are hydrolysed in our bodies to supply us with essential smaller molecules and energy.
- Small molecules within our bodies can undergo condensation polymerisation reactions to be stored as larger molecules.
- Biodiesel can be formed from the esterification reaction between a plant triglyceride and methanol.

9.3

- Atom economy is important for chemists to consider when trying to minimise waste.
- Percentage yield can be calculated after a chemical reaction has been completed.

9.4

- Green chemistry principles help chemists work toward Sustainable Development Goal 12.
- Renewable feedstocks are being used to move away from our reliance on fossil fuels.
- Catalyst development and selection and atom economy work together to minimise waste products.
- Chemical design should aim to increase the efficiency of reactions and decrease the toxicity of products produced.

Key formulas

Percentage yield	$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$
Percentage atom economy	$\text{percentage atom economy} = \frac{\text{molar mass of desired product}}{\text{molar mass of all reactants}} \times 100$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as ‘Confidently’, ‘Mostly’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Mostly	Not really	Revision link
• describe different organic reactions and pathways, including equations, reactants, products reaction conditions and catalysts for: synthesis of primary haloalkanes and primary alcohols by substitution, addition reactions of alkenes, the esterification between an alcohol and a carboxylic acid, hydrolysis of esters, pathways for the synthesis of primary amines and carboxylic acids	☐	☐	☐	Go back to Topic 9.1
• describe different organic reactions and pathways, including equations, reactants, products reaction conditions and catalysts for: transesterification of plant triglycerides using alcohols to produce biodiesel, hydrolytic reactions of proteins, carbohydrates and fats and oils to break down large biomolecules in food to produce smaller molecules, condensation polymerisation reactions to synthesise large biologically important molecules for storage as proteins, starch, glycogen and lipids (fats and oils)	☐	☐	☐	Go back to Topic 9.2
• calculate percentage yield and atom economy of single-step or overall reaction pathways	☐	☐	☐	Go back to Topic 9.3
• describe the advantages of developing chemical processes with a high atom economy, for society and for industry	☐	☐	☐	Go back to Topic 9.3
• describe the sustainability of the production of chemicals, with reference to the green chemistry principles of use of renewable feedstocks, catalysis and designing safer chemicals	☐	☐	☐	Go back to Topic 9.4

Revision questions

Multiple choice

1 Which of the options correctly identifies the bonds formed between the biomolecules?

	Starch	Triglyceride	Protein
A	Glycosidic	Ether	Amide
B	Glycosidic	Ester	Peptide
C	Ester	Amide	Peptide
D	Amide	Glycosidic	Ester

2 Which of the following statements are true of percentage yield?

I Percentage yield is a comparison of actual yield and theoretical yield.

II Percentage yield is used to calculate the most sustainable pathway for a reaction.

III Percentage yield can only be calculated after a reaction is completed.

A I only

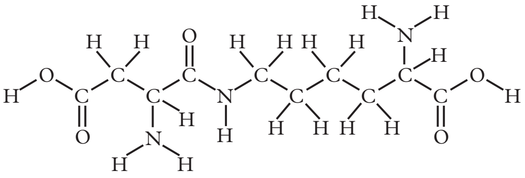
B I and II

C I and III

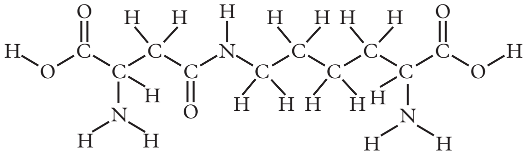
D II and III

3 Identify the correct dipeptide formed from the condensation reaction between lysine and aspartic acid.

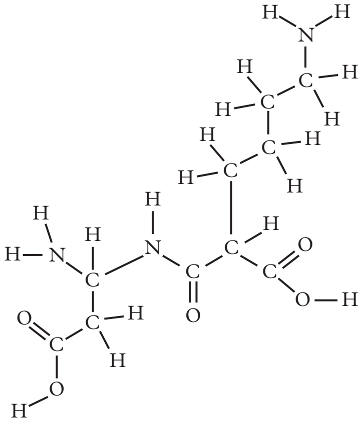
A



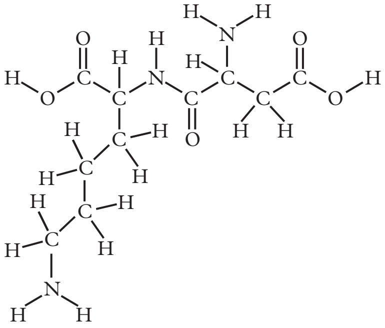
B



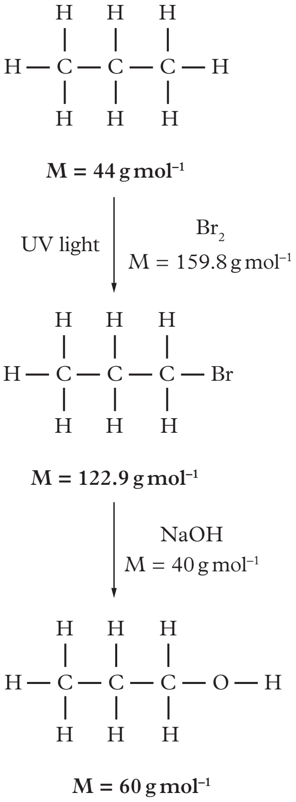
C



D



4 What is the atom economy for the two-step reaction shown?



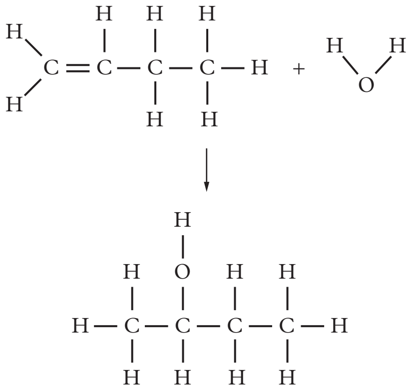
A 16%

B 25%

C 41%

D 75%

5 Which of the following catalysts is required for the reaction shown?



A H_2SO_4

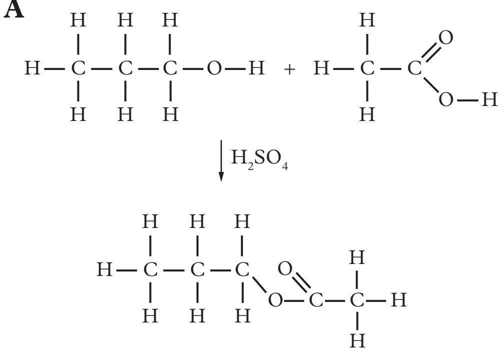
B H_3PO_4

C UV light

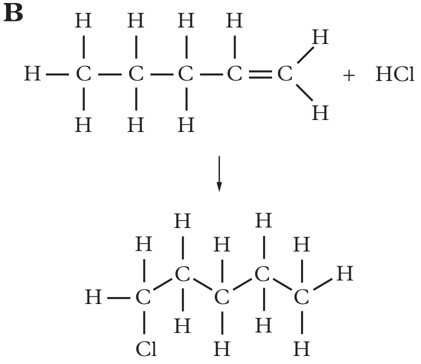
D $H^+/KMnO_4$

6 Identify which of the following reactions will *not* result in the product shown.

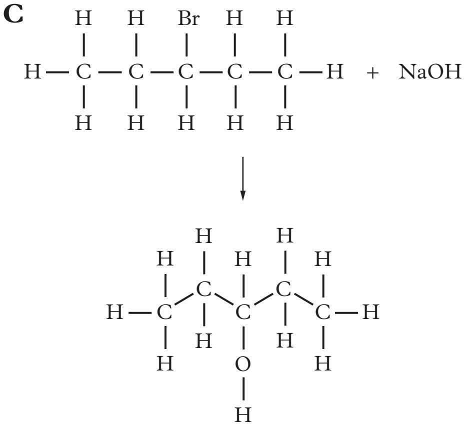
A



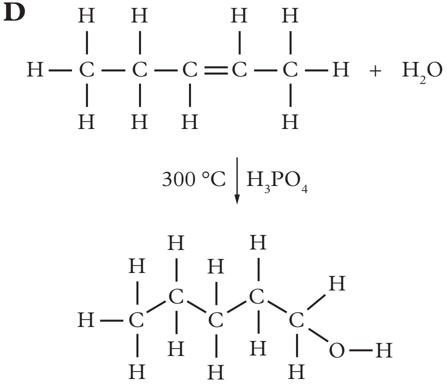
B



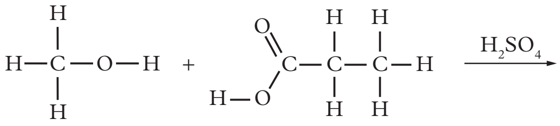
C



D



7 Identify the correct IUPAC name for the product formed in the reaction shown.



A Methyl ethanoate

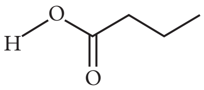
B Ethyl ethanoate

C Propyl methanoate

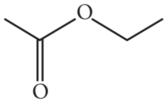
D Methyl propanoate

8 Br_2 was added to an unknown compound. The product was a colourless liquid. Identify which of the compounds below is the unknown compound.

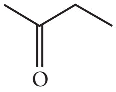
A



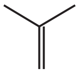
B



C



D



9 Which of the following is *not* an oxidising agent for the oxidation of an alcohol into a carboxylic acid?

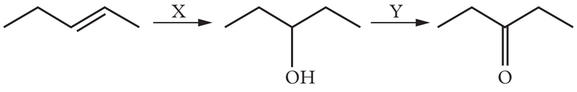
A $H^+/KMnO_4$

B $H^+/Cr_2O_7^{2-}$

C H^+/H_3PO_4

D $H^+/K_2Cr_2O_7$

10 Consider the reaction pathway shown.



Compound A **Compound B** **Ketone**

Identify the correct reagents and conditions for X and Y and the IUPAC name of compound B.

	X	Y	Compound B
A	H_2O/H_3PO_4	$Cr_2O_7^{2-}/H^+$	Pentan-3-ol
B	NaOH	$KMnO_4/H^+$	Hexan-3-ol
C	NaOH	$KMnO_4/H^+$	Pentan-3-ol
D	H_2O/H_3PO_4	$Cr_2O_7^{2-}/H^+$	Hexan-3-ol

Short answer

Describe and explain

11 Identify the reagents and conditions for the following reactions.

- **a** Ethanol → ethanoic acid
- **b** Hexene + H_2 → hexane
- **c** Butane + Cl_2 → 1-chlorobutane + NH_3 → butan-1-amine
- **d** Pent-2-ene → pentan-3-ol

12 Describe the reaction types that occurred in Question 11.

13 The molecule ethanol can be synthesised from biomass or a fossil fuel.

- **a** Construct the pathway for the production of ethanol from biomass and from fossil fuel.
- **b** Explain how the pathways differ.
- **c** Explain if there is a difference between the molecules bioethanol and ethanol.

14 Methanol cannot be directly synthesised from methane. Explain why this is true and create a pathway to synthesise methanol.

15 Explain why pentan-2-ol cannot be oxidised into a carboxylic acid.

16 Describe three ways in which you could increase the percentage yield in a reaction.

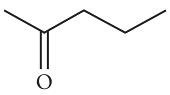
17 Describe the types of reactions that will have high atom economies. Use examples to justify your response.

18 Explain why atom economy is important to consider when designing reaction pathways.

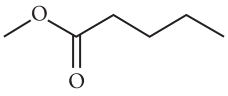
Apply, analyse and compare

19 Consider the three molecules below.

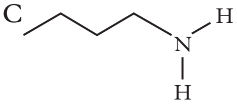
A



B

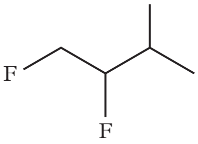


C

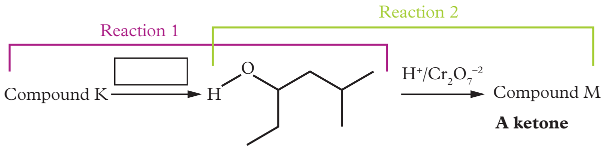


- **a** Complete the oxidation reaction for molecule A from its secondary alcohol.
- **b** Analyse molecule B.
  - **i** Complete the reaction pathway to synthesise compound B from an alkane and an alkene.
  - **ii** Calculate the atom economy of your pathway in part **b i**.
  - **iii** Explain if you could increase the atom economy at any step of your pathway.
- **c** Complete the pathway for molecule C if you started with an alkene.

20 Analyse the compound shown and identify the reactants used to synthesise it.



21 Analyse the reaction pathway below and answer the following questions.



- **a** Draw the full structural formula for compounds K and M.
- **b** Identify the catalyst and conditions for reaction 1.
- **c** Give the IUPAC name for the alcohol in this reaction.
- **d** Identify the types of reactions that are occurring in this pathway.
- **e** Calculate the atom economy for each reaction.

22 Compare the rate of hydrolysis of amylose and amylopectin. Explain which of the two starches will hydrolyse faster and why.

23 The two reactions on the next page form hexan-1-ol. Calculate and compare their atom economies. Identify the more efficient/sustainable pathway.

Pathway 1:

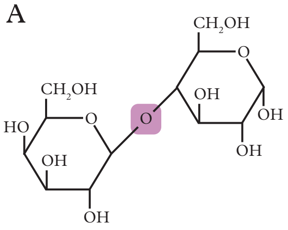
$CH_3CH_2CH_2CH_2CH_2CH_3 + Cl_2 \rightarrow$
$CH_3CH_2CH_2CH_2CH_2CH_2Cl + NaOH \rightarrow$
$CH_3CH_2CH_2CH_2CH_2CH_2OH$

Pathway 2:

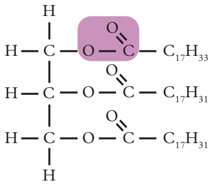
$CH_3CH_2CH_2CH_2CH{=}CH_2 + HCl \rightarrow$
$CH_3CH_2CH_2CH_2CH_2CH_2Cl + NaOH \rightarrow$
$CH_3CH_2CH_2CH_2CH_2CH_2OH$

24 Three biomolecules are shown. Analyse their structures and answer the following questions.

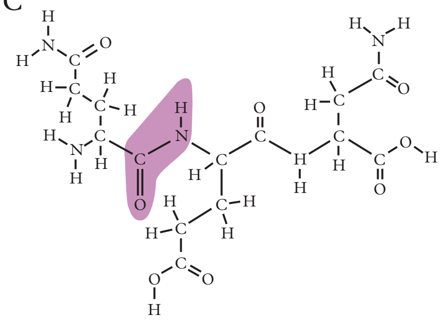
A



B



C

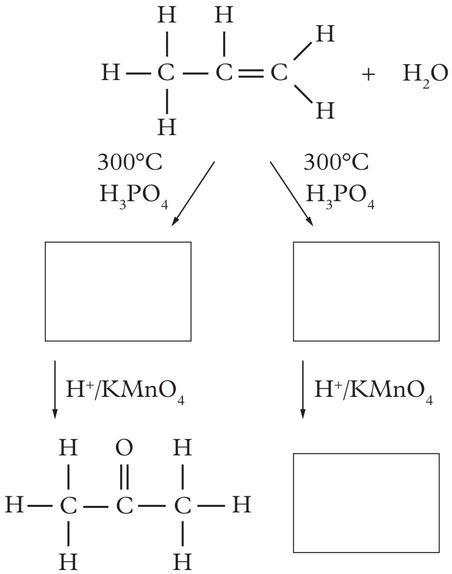


- **a** Identify each of the functional groups highlighted purple on molecules A–C.
- **b** Explain what sort of biomolecules A–C are.
- **c** Identify the names of the monomers that make up each of the biomolecules A–C.
- **d** Complete the transesterification reaction that molecule B would go through to synthesise biodiesel.

25 Propene can be reacted with water in an addition reaction to form two different isomers as seen in the reaction shown. Both isomers can then undergo oxidation reactions.

- **a** Complete the pathways below.
- **b** Apply IUPAC naming rules to the three products you have drawn.

- **c** Calculate and compare the atom economies of the two reactions.



26 Freon, CF_2Cl_2, is produced from the reaction shown.

$$CCl_4 + 2HF \rightarrow CF_2Cl_2 + 2HCl$$

- **a** Identify the type of reaction that produces freon.
- **b** Calculate the percentage yield when a reaction produces 25.0 g of freon, from 65.8 g of CCl_4 and excess HF.

The reaction to produce freon is undertaken again, but this time, 43.0 g of CCl_4 is reacted with 7.30 g of HF to produce 12.6 g of freon.

- **c** Explain which reactant is the limiting reagent in this reaction.
- **d** Calculate the percentage yield for this reaction.

FIGURE 1 Freon gas is used in air conditioners.

27 **a** Calculate the atom economy of the production of biodiesel from a plant triglyceride with three linoleic acid residues using a KOH catalyst.

b If you used a lipase-enzyme catalyst instead of KOH, explain if the atom economy will change or not.

28 Analyse the structures of amylopectin and glycogen and answer the questions.

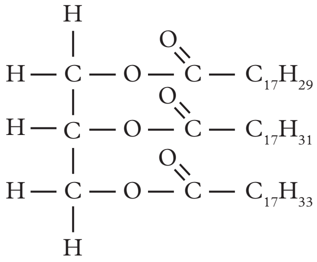
a Identify the monomers that make up amylopectin and glycogen.

b Compare the structure of amylopectin with that of glycogen. Explain how they are similar and how they are different.

c Describe the hydrolysis of amylopectin.

d Describe the condensation reaction that would take place to synthesise glycogen.

29 Analyse the triglyceride below and answer the questions.



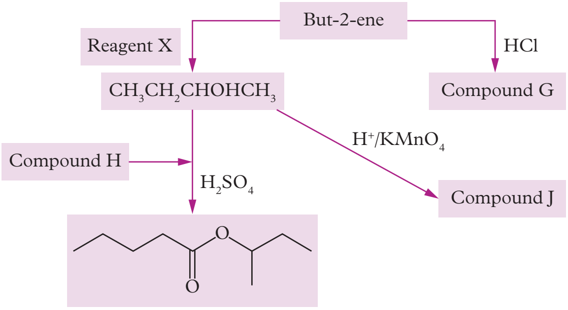
a Name the four molecules that would react to form this triglyceride.

b Suggest if it would more likely be a plant or animal triglyceride. Explain your reasoning.

c Discuss if it would more likely be a liquid or solid at room temperature.

d Draw the transesterification reaction pathway for this triglyceride to become biodiesel.

30 Analyse the pathway and answer the questions.



Ester: 1-methylpropyl pentanoate

a Identify reagent X.

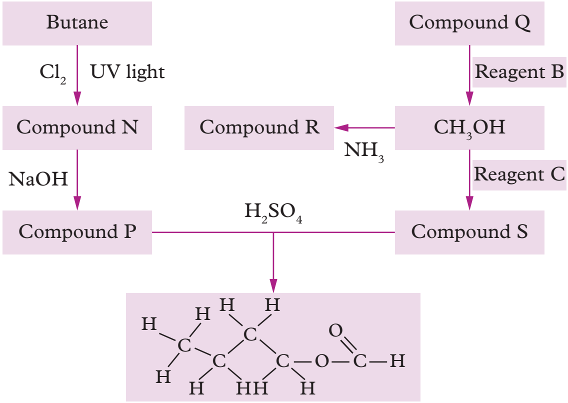
b Draw the structural formula for compound H.

c Name and draw the skeletal structure for compound G.

d Draw the structural formula for compound J.

e Identify the reaction occurring to form compound J.

31 Analyse the pathway and answer the following questions.



a Identify the names of and draw the structural formulas for compounds Q, R and S.

b Identify the names of compounds N and P and draw the sturctural formulas for each.

c Identify reagents B and C.

d Name the ester formed in this reaction.

e Identify the reaction types for each step in the pathway.

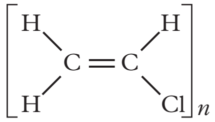
f Calculate the atom economy of the reaction from butane to compound P.

Design and discuss

32 Discuss the advantages and disadvantages of using an enzyme to catalyse the reaction of plant triglycerides into biodiesel over the commonly used inorganic hydroxide catalyst.

33 Design a pathway to produce an ester starting with only butane.

34 Consider the monomer that makes up the plastic polyvinylchloride PVC.



Chloroethene or vinyl chloride

a Discuss whether you could synthesise this monomer from a renewable feedstock.

b Complete the reaction pathway to create polyethylene from a renewable feedstock.

35 A chemist wants to create a new organic surfactant. Discuss the green chemistry principles they would have to consider before starting the process.

36 a Design a pathway with a high atom economy for the synthesis of ethyl butanoate. You must start with one alkane and one alkene.

b Justify each step of your pathway and why it has the highest percentage atom economy.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

UNIT 4

AREA OF STUDY 1

Checkpoint

Multiple choice

Question 1

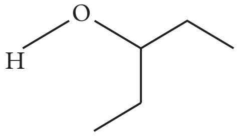
Which of the following is not a possible isomer of $C_5H_{10}O_2$?

A 3-methylbutanoic acid
B 2-methylpropanoic acid
C butyl methanoate
D propyl ethanoate

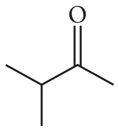
Question 2

Which of the following is not a possible isomer of $C_5H_{10}O$?

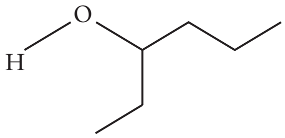
A



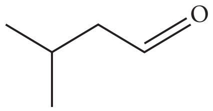
B



C



D



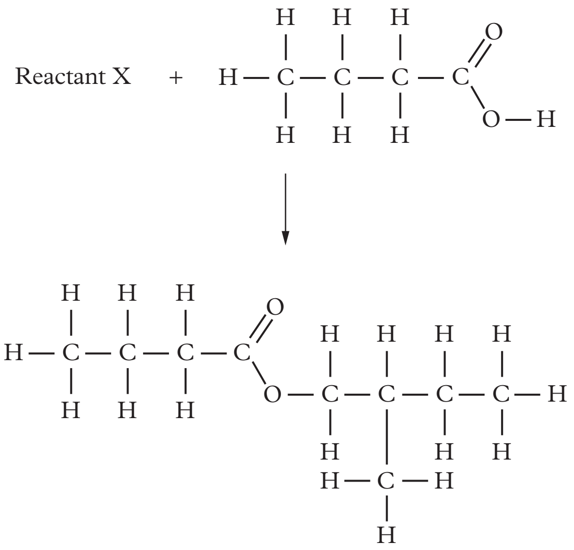
Question 3

Identify which of the following isomers of $C_6H_{12}O$ will have the lowest boiling point.

A $CH_3CH_2CH_2COCH_2CH_3$
B $CH_3CH_2CH_2CH_2CH_2CH_2OH$
C $CH_3CH_2CH_2CH_2CH_2CHO$
D $(CH_3)_2CHCOCH_2CH_3$

Question 4

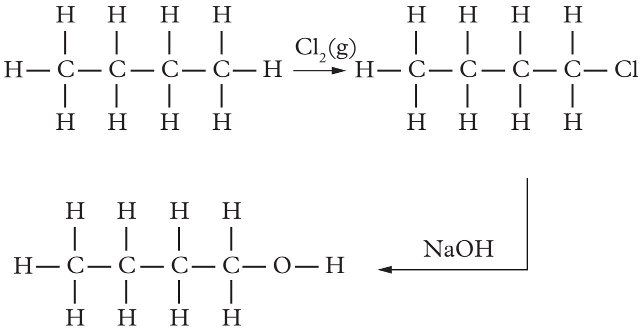
Identify the correct IUPAC name for reactant X in the following reaction:



A 2-methylbutan-1-ol
B 3- methylbutan-4-ol
C butanoic acid
D butanol

Question 5

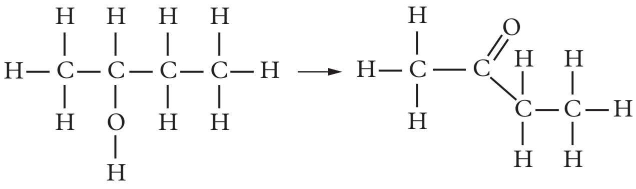
What is the atom economy for the two-step reaction pathway to form butan-1-ol from butane:



A 28%
B 31%
C 36%
D 44%

Question 6

Which of the following catalysts is required for the reaction shown?



A H_2SO_4

B H_3PO_4

C UV light

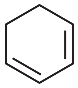
D $H^+/KMnO_4$

Question 7

From the following list, identify which molecule(s) will have a degree of unsaturation of 2:

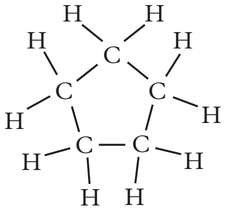
I $C_4H_7NH_2$

II



III C_3H_5Cl

IV



A Molecule II only

B Molecules I, II and IV

C Molecules I and III

D None of the molecules

Question 8

Which of the following statements is incorrect?

A As electronegativity increases, bond energy increases.

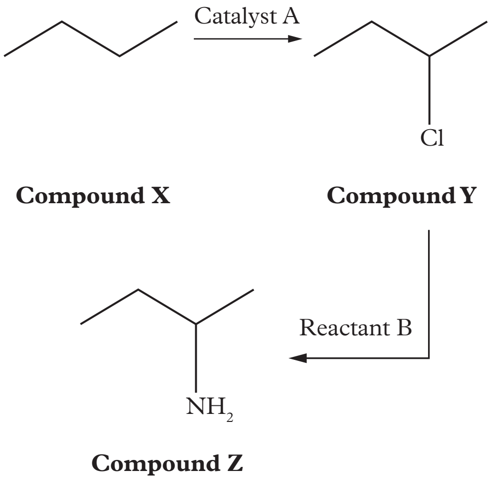
B Bond length increases as the number of bonds increase.

C The energy of a C–halogen bond decreases as the size of the halogen increases.

D As bond energy decreases, bond stability also decreases.

Question 9

A reaction pathway is shown.

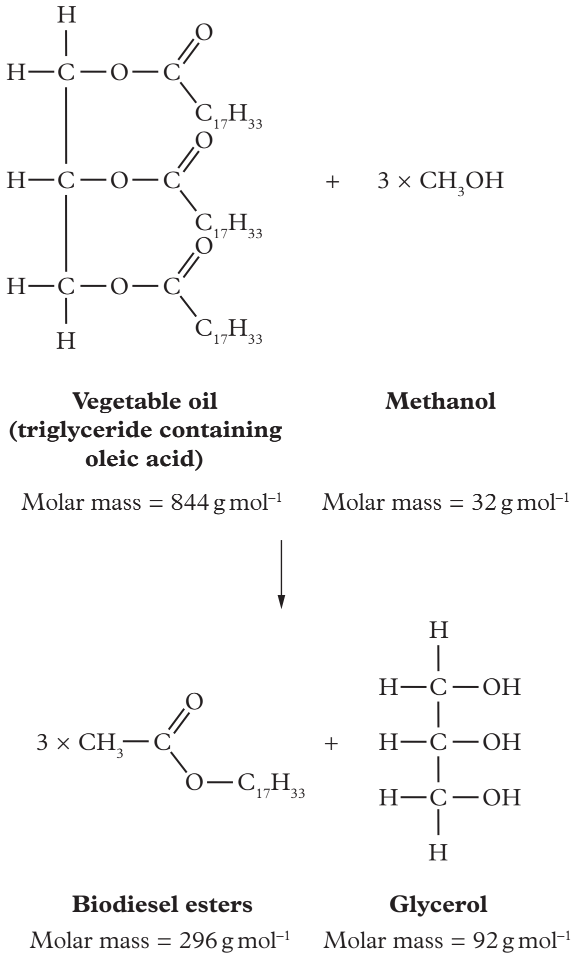


Identify Catalyst A, Reactant B and Compound Z in the table.

	Catalyst A	Reactant B	Compound Z
A	NaCl	$H^+/KMnO_4$	propan-2-amine
B	UV/heat	NH_2	butanamine
C	UV/heat	NH_3	butan-2-amine
D	NaCl	H_3PO_4	propanamine

Question 10

Students complete a reaction to synthesise biodiesel from vegetable oil containing primarily triglycerides of oleic acid. They react 90.0 g of oil with excess methanol. This produces 83.4 g of biodiesel after the glycerol waste is removed. The reaction pathway is shown below.



Vegetable oil (triglyceride containing oleic acid)

Molar mass = 844 g mol^{-1}

Methanol

Molar mass = 32 g mol^{-1}

Biodiesel esters

Molar mass = 296 g mol^{-1}

Glycerol

Molar mass = 92 g mol^{-1}

Calculate the percentage yield from the reaction that the students performed.

A 36.1%

B 88.1%

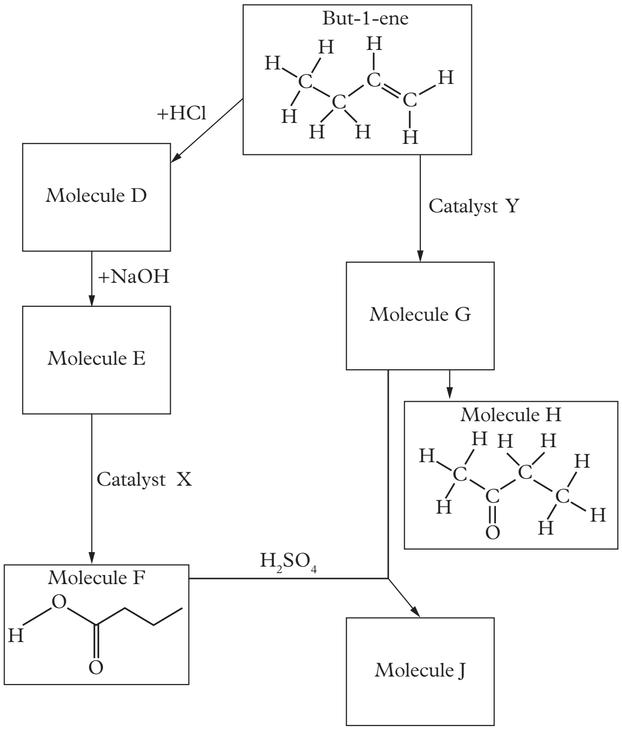
C 92.7%

D It cannot be calculated due to insufficient information.

Short answer

Question 1 (10 marks)

The following pathway shows one way in which but-1-ene can be synthesised to form a branched ester, called molecule J.



a Write the semi-structural formula of Molecule D. 1 mark

b Identify the type of reaction occurring to form Molecule E from Molecule D. 1 mark

c Write the molecular formula for Molecule E. 1 mark

d Identify Catalyst X. 1 mark

e Apply the IUPAC naming rules to name Molecule F. 1 mark

f Identify Catalyst Y and the reaction conditions required to form Molecule G from but-1-ene. 2 marks

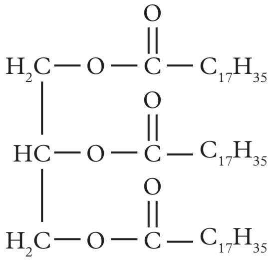
g Draw the skeletal structure of Molecule G. 1 mark

h Apply the IUPAC naming rules to name Molecule E. 1 mark

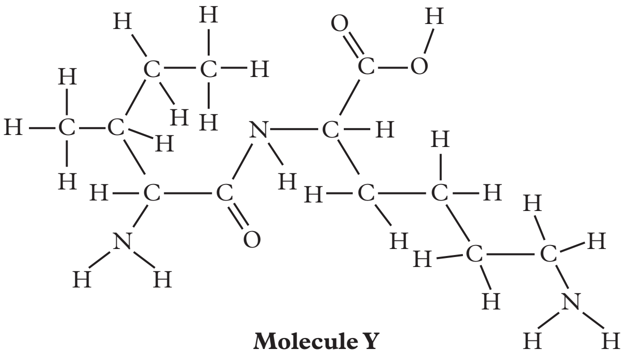
i Draw the full structural formula of Molecule J. 1 mark

Question 2 (11 marks)

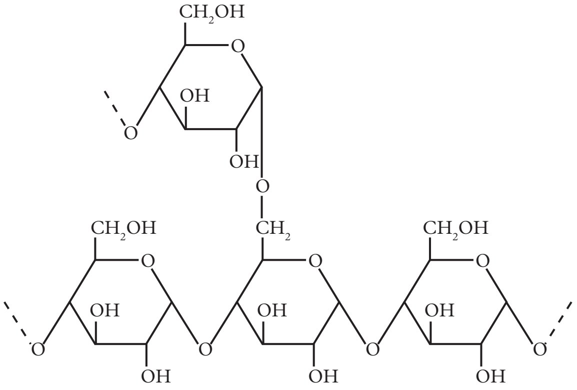
Three biomolecules are shown.



Molecule X



Molecule Y



Molecule Z

a Identify the types of biomolecules that molecules X, Y and Z are. 3 marks

b One of the biomolecules is formed when a condensation reaction occurs between amino acids.

- **i** Identify which molecule this is. 1 mark
- **ii** On the diagram, circle the bond that is formed when this condensation reaction between amino acids occurs. 1 mark
- **iii** Identify the name of this bond. 1 mark
- **iv** Draw the structural formula of one of the amino acids that makes up this molecule. 1 mark

c Molecule X can be broken down into fatty acids in the body. Draw the fatty acid that would be formed from this molecule. 1 mark

d One of the molecules can undergo a reaction to form monomers that have the molecular formula $C_6H_{12}O_6$.

- **i** Identify which of the molecules this is. 1 mark
- **ii** Identify the name of the bonds that are broken during this reaction. 1 mark
- **iii** Name the type of reaction that takes place to form these monomers. 1 mark

Question 3 (11 marks)

Some students decide to find out which oil makes the biofuel with the highest energy content.

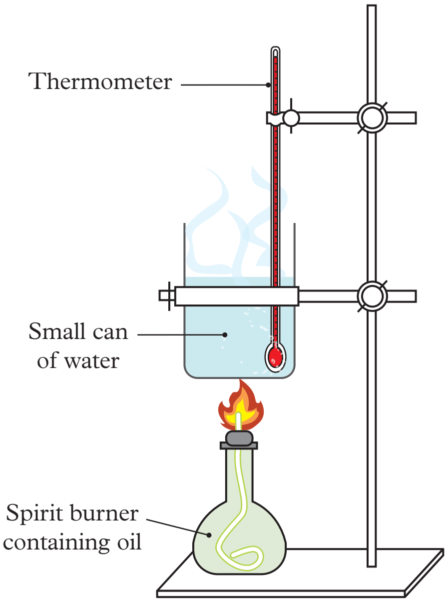
They use three different cooking oils to synthesise their biodiesel methyl esters: canola oil, olive oil and peanut oil. Olive oil contains the fatty acids stearic acid, palmitic acid, linoleic acid, oleic acid and linolenic acid.

a Draw the triglyceride that would be formed from the synthesis of three linoleic acids and a glycerol. 2 marks

To make the biodiesel from the triglyceride, methanol must be added.

b Draw the pathway of the transesterification of the plant triglyceride you drew in part **a** to produce biodiesel molecules. 3 marks

After synthesising the peanut, canola and vegetable oils into biodiesel esters, the students must test the energy content of each. They set up a spirit burner as shown in the figure.



The students measure the initial and final mass of the oil used to heat 100 g of water in the can, and the temperature change of the water. Between each oil tested, they change the water.

The data collected by the students is shown in the table.

Oil	Initial mass of oil (g)	Final mass of the oil (g)	Change in temperature of 100 g of water
Canola	27.25	26.95	14.2°C
Peanut	30.10	29.70	11.6°C
Olive	25.20	24.95	15.1°C

c Calculate the energy content (kJ/g) of the olive oil. 2 marks

d Compare the energy content (kJ/g) of the three oils and write an answer to the student's investigation question. 2 marks

e In their discussion, the students claim the investigation is related to the green chemistry principles. Identify the green chemistry principle that this is likely to refer to and explain how the investigation relates to this principle. 2 marks

Question 4 (8 marks)

The table contains four molecules and their boiling points.

Molecule	Boiling point (°C)
Pentane	36.1
Pentan-1-ol	138
Pentanal	103
Pentanoic acid	186

a Identify the strongest intermolecular forces that exist between pentanoic acid molecules. 1 mark

b The molar mass of pentanal is 86 $g\,mol^{-1}$ and the molar mass of pentan-1-ol is 88 $g\,mol^{-1}$, yet the difference in boiling point is 35°C. Explain why pentan-1-ol has a significantly higher boiling point than a molecule of a similar molar mass. 2 marks

c Draw a branched isomer of pentan-1-ol. 1 mark

d Predict the boiling point of the isomer of pentan-1-ol you drew and compare this with the boiling point of pentan-1-ol from the table above. Explain your reasoning. 2 marks

e Discuss the trend in the viscosity of the four molecules in the table. 2 marks

TOTAL MARKS

/50 marks

CHAPTER 10

Laboratory analysis of organic compounds

KEY KNOWLEDGE

- qualitative tests for the presence of carbon–carbon double bonds, hydroxyl and carboxyl functional groups
- applications of laboratory analysis techniques in verifying components and purity of consumer products for society, including melting point determination and distillation (simple and fractional)
- measurement of the degree of unsaturation of compounds using iodine
- volumetric analysis, including calculations of excess and limiting reactants using redox titrations (excluding back titrations).

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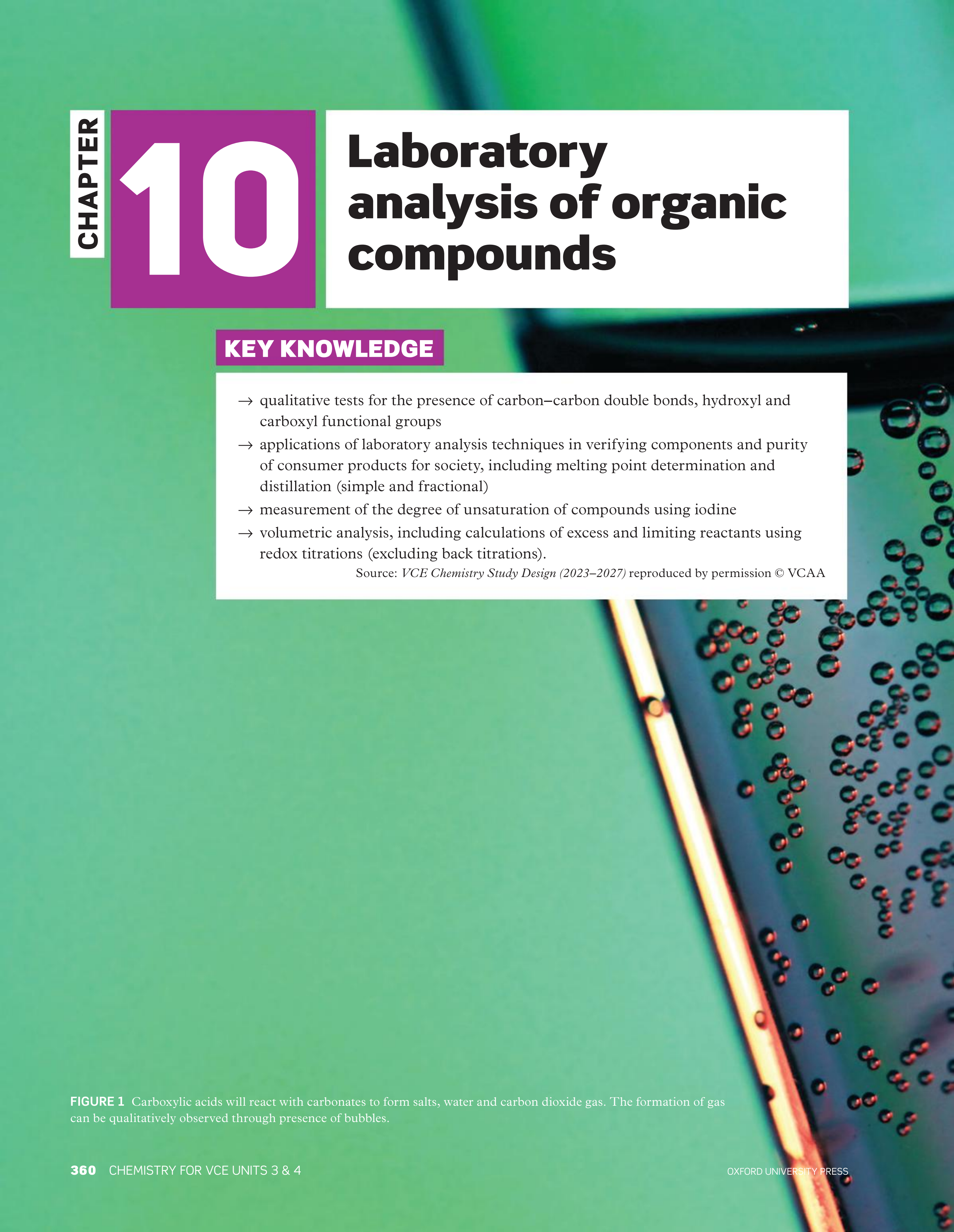
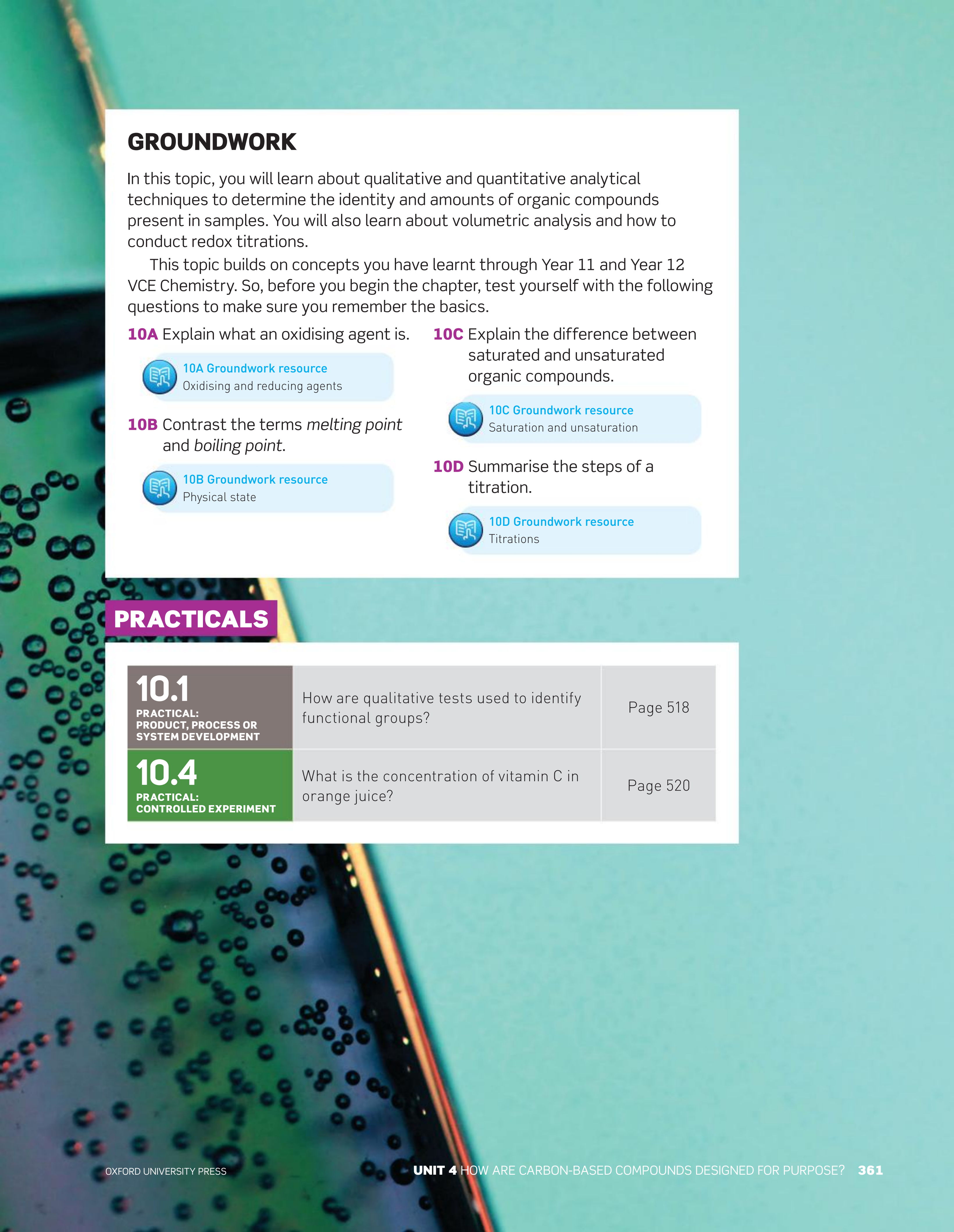


FIGURE 1 Carboxylic acids will react with carbonates to form salts, water and carbon dioxide gas. The formation of gas can be qualitatively observed through presence of bubbles.



GROUNDWORK

In this topic, you will learn about qualitative and quantitative analytical techniques to determine the identity and amounts of organic compounds present in samples. You will also learn about volumetric analysis and how to conduct redox titrations.

This topic builds on concepts you have learnt through Year 11 and Year 12 VCE Chemistry. So, before you begin the chapter, test yourself with the following questions to make sure you remember the basics.

10A Explain what an oxidising agent is.

10A Groundwork resource
Oxidising and reducing agents

10B Contrast the terms *melting point* and *boiling point*.

10B Groundwork resource
Physical state

10C Explain the difference between saturated and unsaturated organic compounds.

10C Groundwork resource
Saturation and unsaturation

10D Summarise the steps of a titration.

10D Groundwork resource
Titrations

PRACTICALS

10.1 **PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT**	How are qualitative tests used to identify functional groups?	Page 518
10.4 **PRACTICAL: CONTROLLED EXPERIMENT**	What is the concentration of vitamin C in orange juice?	Page 520

10.1 Qualitative chemical tests

KEY IDEAS

In this topic, you will learn that:

- carbon–carbon double bonds can be identified using addition reactions by breaking the double bond using bromine
- the position of hydroxyl groups (–OH) can be identified as being primary, secondary, or tertiary, using oxidation reactions
- the presence of carboxyl groups (–COOH) can be identified through acid–base and esterification reactions.

qualitative testing
tests that gather data that is not measurable; often based on observations or descriptions

Now that you have learnt about the categorisation and synthesis of organic compounds, you will look at how they are analysed and used. This topic focusses on **qualitative testing** of organic compounds or, in other words, their identification or characterisation.

Qualitative tests for functional groups

Qualitative tests can determine the presence of an acid or base, organic functional group, carbon–carbon double bond or the identity of a gas. These characteristics are indicated by a colour change, the formation of a gas or a change in texture/precipitate.

Qualitative testing for double bonds

If you recall from Chapter 8, a saturated hydrocarbon (e.g. an alkane) has carbon–carbon single bonds only, whereas an unsaturated hydrocarbon (e.g. an alkene) has one or more carbon–carbon double bonds. We can test for the presence of carbon–carbon double bonds in a compound.

You will look at two ways to do this: the bromine test and the iodine test. The bromine test is covered here, and the iodine test is covered in Topic 10.4.

FIGURE 1 The changes that occur in qualitative chemical tests, such as colour change, are observational.

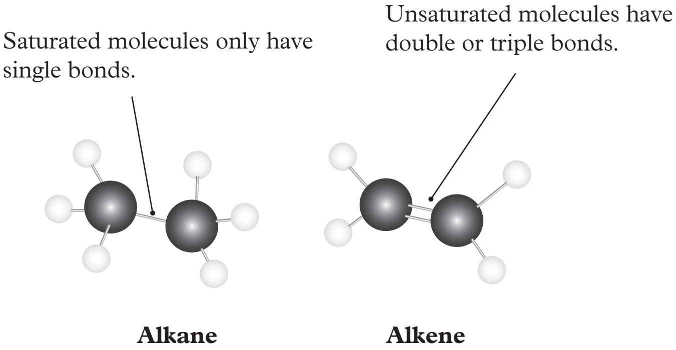


FIGURE 2 Saturated versus unsaturated molecules

Both of these tests for unsaturation use addition reactions. The reactants combine at the double bond to form one product. Reactions are chosen based on their ability to produce a colour change when the carbon–carbon double bonds are broken.

10.1
PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT
How are qualitative tests used to identify functional groups? Go to page **518**.

The bromine test

Bromine gas (Br_2) is a chemical compound that is non-polar and deep red-orange in colour. It is dissolved in water to form a bromine–water mixture, which is deep orange in colour (Figure 3). Bromine can break the double bonds of unsaturated organic compounds and form new covalent bonds.

When the bond between bromine atoms breaks, bromine loses its colour. Therefore, if bromine has reacted with an unsaturated organic compound, the solution will be colourless. Bromine will not react with a saturated substance. In this case, it remains intact, and the solution will be coloured.

FIGURE 3 Bromine water is deep orange in colour.

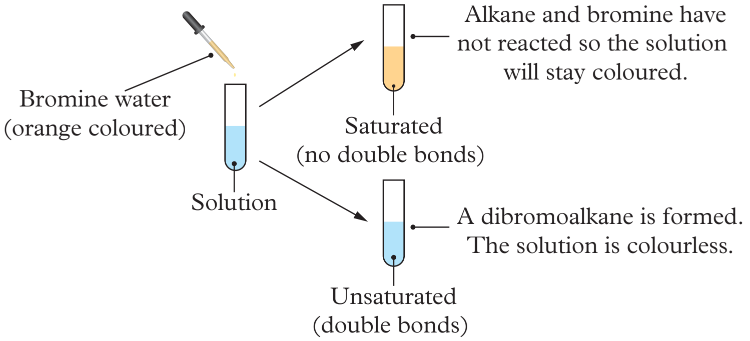


FIGURE 4 When bromine water is added to an alkane, the solution will be coloured. When added to an alkene, the solution will be clear.

In a **bromine test**, bromine water is added to a test tube with the hydrocarbon. Since hydrocarbon compounds are non-polar and water molecules in bromine water are polar, the solutions form layers that do not initially mix. The hydrocarbon is less dense than the bromine water, so it forms the top layer in the test tube. The test tube is shaken to increase the interaction between the bromine in the bromine water and the hydrocarbon, then allowed to separate into layers again. We look at the colour of the bottom layer to determine whether our hydrocarbon is saturated or unsaturated.

bromine test
a qualitative test in which bromine water is added to a hydrocarbon sample to test for carbon–carbon double bonds; if the bromine changes from orange to colourless, the hydrocarbon is unsaturated

If the hydrocarbon is saturated, it will not react. The bromine will remain in the water layer, so the bottom layer will remain coloured (Figure 5a). If the hydrocarbon is unsaturated, bromine will react. The bromine will be consumed in the reaction and the bottom layer will become colourless (Figure 5b).

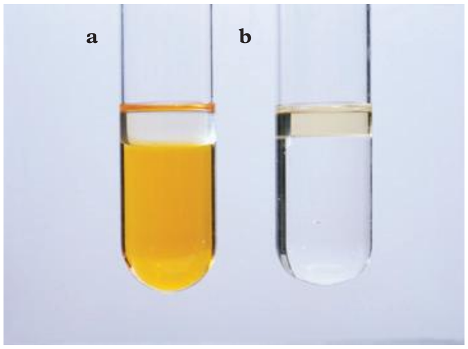


FIGURE 5 The reaction between bromine water and **a** an alkane and **b** an alkene.

Let's look at an example. Bromine will break the double bond between the two carbon atoms in ethene, allowing a bromine atom to covalently bond to each of the two carbon atoms (Figure 6a). A similar observation is made for but-2-ene, where the double bond is between the second and third carbon atom. When the double bond breaks, one bromine atom bonds to the second carbon atom and one bonds to the third carbon atom (Figure 6b). Would you expect the test tubes to look more like Figure 5a or 5b?

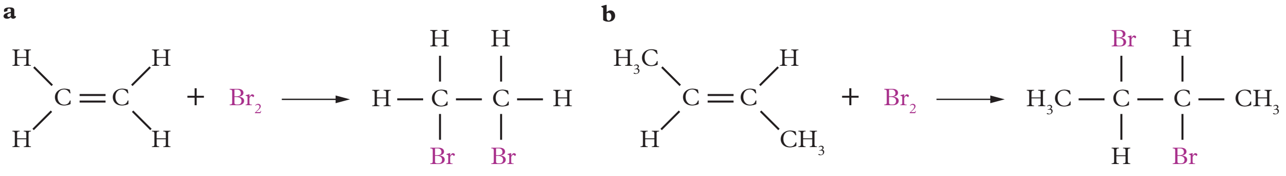


FIGURE 6 The addition reaction between bromine and **a** ethene and **b** but-2-ene

Qualitative testing for hydroxyl groups

Alcohols contain hydroxyl (OH) functional groups, which can be oxidised to form aldehydes, ketones or carboxylic acids. The carbon attached to the hydroxyl group loses electrons and its oxidation number increases. Whether a reaction occurs and the product it forms depends on whether the alcohol is primary, secondary or tertiary in structure (Figure 7).

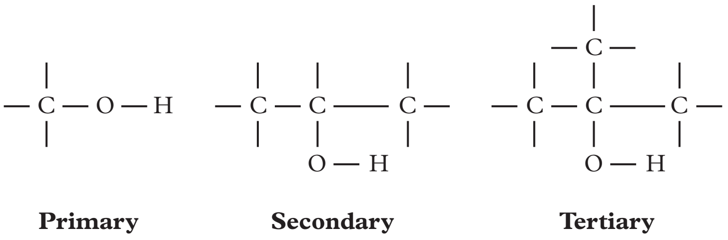
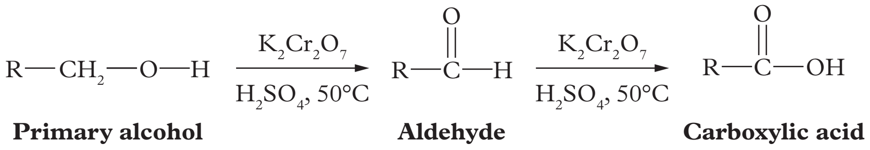


FIGURE 7 The structure of primary, secondary and tertiary alcohols

Oxidation of alcohols

The oxidation of alcohols requires an oxidising agent – a chemical that will cause oxidation but itself undergoes reduction. In qualitative tests, the oxidising agent must be a compound that will change colour when it undergoes reduction. Dichromate and permanganate, in acidic conditions, are two commonly used coloured oxidising agents for testing alcohols.

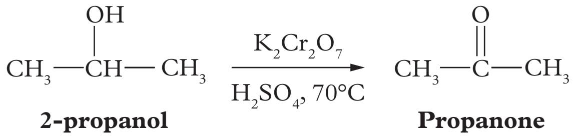
Primary alcohols undergo partial oxidation to form aldehydes, and complete oxidation to form carboxylic acids.



$$R{-}CH_2{-}O{-}H \xrightarrow[H_2SO_4,\ 50°C]{K_2Cr_2O_7} R{-}\overset{\overset{O}{\|}}{C}{-}H \xrightarrow[H_2SO_4,\ 50°C]{K_2Cr_2O_7} R{-}\overset{\overset{O}{\|}}{C}{-}OH$$

FIGURE 8 A primary alcohol is first oxidised to an aldehyde and then to a carboxylic acid.

Secondary alcohols can be oxidised to form ketones (containing carbonyl groups).



$$CH_3{-}\overset{\overset{OH}{|}}{CH}{-}CH_3 \xrightarrow[H_2SO_4,\ 70°C]{K_2Cr_2O_7} CH_3{-}\overset{\overset{O}{\|}}{C}{-}CH_3$$

FIGURE 9 A secondary alcohol oxidised to a ketone

Tertiary alcohols cannot be oxidised because they have no hydrogen atoms available to lose. A colour change will indicate that an alcohol has a primary or secondary structure. No colour change indicates that the alcohol is tertiary.

Figure 10 shows the use of dichromate ($Cr_2O_7^{2-}$), which has an orange/yellow colour from the presence of chromium. Once reduced to Cr^{3+}, the orange/yellow colour changes to a bluish green.

$$\underset{\text{yellow}}{Cr_2O_7^{2-}(aq)} + 14H^+(aq) + 6e^- \rightarrow \underset{\text{green}}{2Cr^{3+}(aq)} + 7H_2O(l)$$

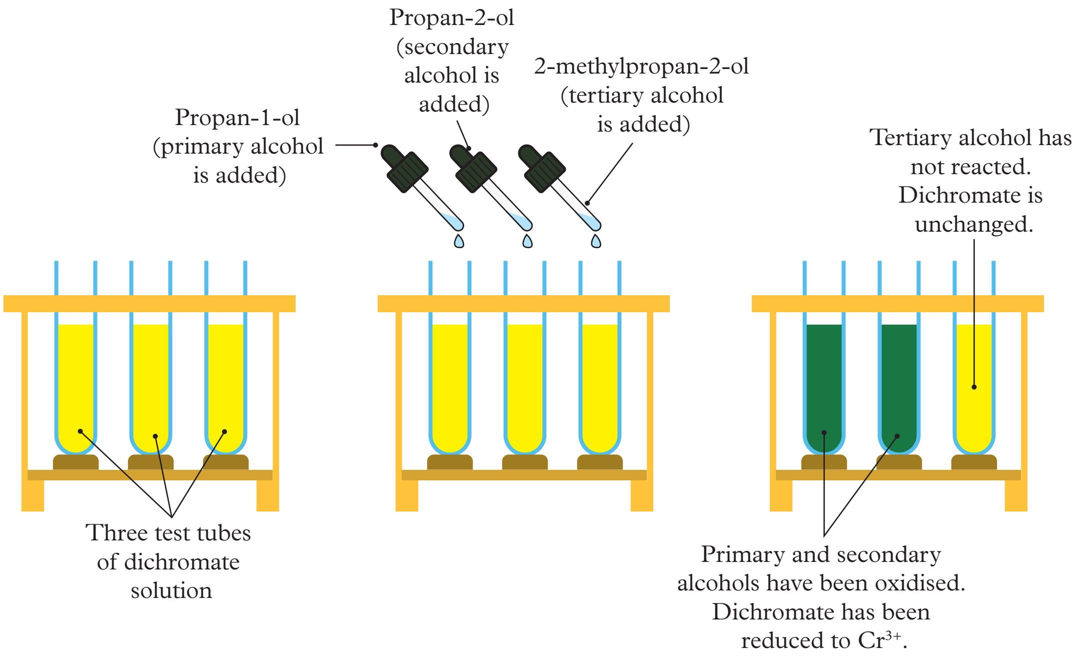


FIGURE 10 The qualitative testing of propan-1-ol, propan-2-ol and 2-methyl-propan-2-ol using a dichromate solution

Figure 11 shows the use of permanganate (MnO_4^-), which has a purple colour from the presence of manganese. When permanganate is reduced to Mn^{2+}, it loses its purple colour and becomes clear.

$$\underset{\text{purple}}{MnO_4^-(aq)} + 8H^+(aq) + 5e^- \rightarrow \underset{\text{clear}}{Mn^{2+}(aq)} + 4H_2O(l)$$

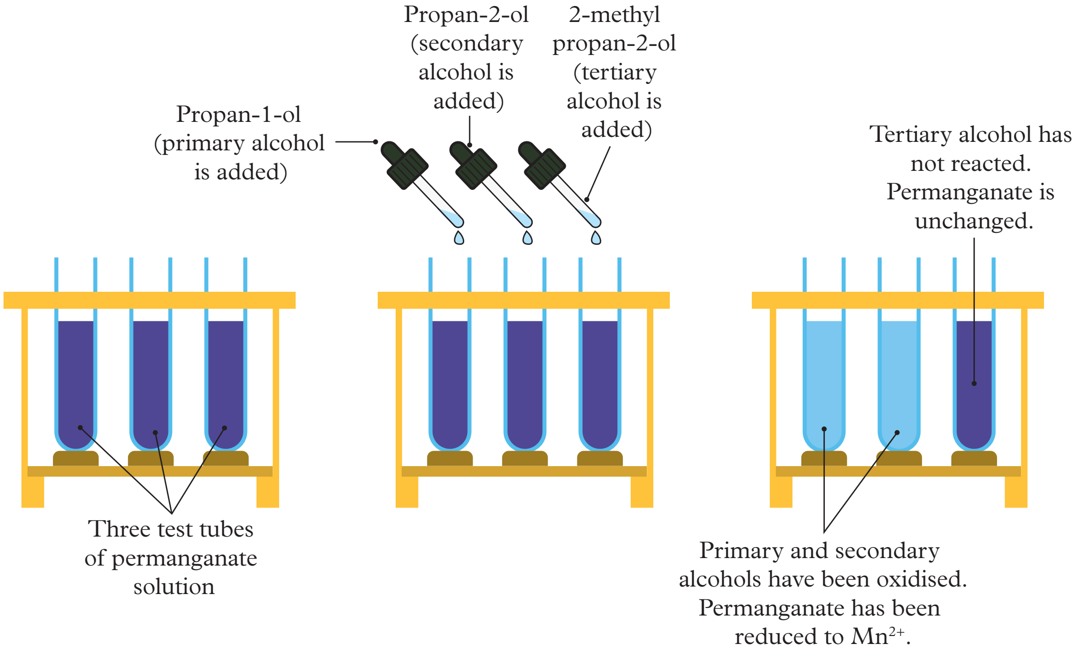


FIGURE 11 The qualitative testing of propan-1-ol, propan-2-ol and 2-methyl-propan-2-ol using a permanganate solution

10.1 Real-world chemistry
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Explore the use of the coloured permanganate solution in breathalysers in Real-world chemistry 10.1.

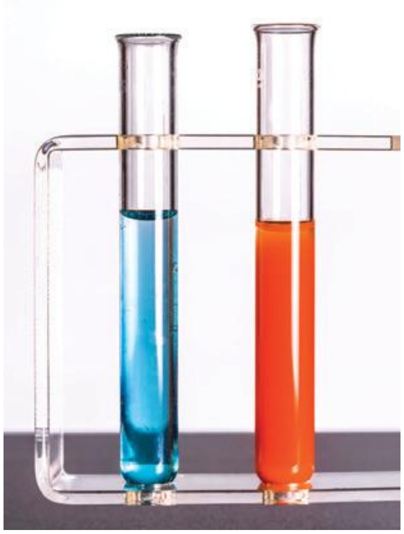
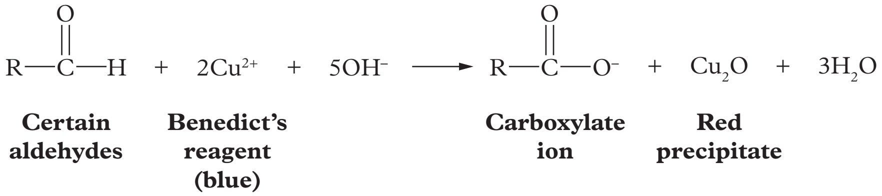


FIGURE 12 The colour of Benedict's solution changes from blue to red when reacted with an aldehyde.

Benedict's test

After using an oxidising agent to confirm the presence of a primary or a secondary alcohol, the alcohols can be distinguished from each other by testing to see if they are an aldehyde or a ketone. This is achieved through a further round of oxidation using **Benedict's solution**. Benedict's solution contains Cu^{2+} ions, which gives the solution a blue colour; these copper ions will react with aldehydes.

In this reaction, the copper ions are reduced to copper oxide and the aldehyde is oxidised to a carboxylic acid. The colour of the solution changes from blue to red. Ketones will not react and the colour will remain unchanged, because they cannot undergo further oxidation.



$$R{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}H + 2Cu^{2+} + 5OH^- \longrightarrow R{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}O^- + Cu_2O + 3H_2O$$

FIGURE 13 Aldehydes will react to Benedict's solution to form carboxylic acids.

Benedict's solution
a mixture of copper(II) sulfate pentahydrate, sodium carbonate and sodium citrate

Study tip

Oxidation reactions can also be used to test for aldehydes, as they will oxidise into carboxylic acids, whereas ketones will not.

Qualitative testing for carboxylic acids

Carboxylic acids are weak organic acids that can react with bases in neutralisation reactions. Carboxylic acids cannot be oxidised any further than they already are, so they are typically tested for by using acid–base reactions.

Acid–base reactions

Carbonates are moderately strong bases. They react with carboxylic acids to form salts, gaseous carbon dioxide and water. In the reaction between ethanoic acid and sodium hydrogen carbonate, the products formed are sodium ethanoate, carbon dioxide gas and water.

$$CH_3COOH(aq) + NaHCO_3(aq) \rightarrow CH_3COONa(aq) + CO_2(g) + H_2O(l)$$

FIGURE 14 The reaction between sodium carbonate and ethanoic acid produces visible carbon dioxide gas bubbles.

The gas formed from the reaction is observed as bubbles. Higher concentrations of carboxylic acids will produce a larger amount (in mol) and mass of carbon dioxide gas. This will produce more bubbles in the reaction mixture. Reactions between calcium carbonate ($CaCO_3$) or sodium carbonate (Na_2CO_3) and carboxylic acids produce similar results.

Indicators

Carboxylic acids ionise in water to form the hydronium ion:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

As carboxylic acids are weak acids, the pH of the solution can be measured. Weak acids produce a solution with a pH below 7 and generally above 3. The indicators in Figure 15 can be used to identify their presence.

Indicator	pH range
Methyl violet	0.15–3.2
Thymol blue	1.2–2.8 and 8.0–9.6
Methyl orange	3.2–4.4
Methyl red	4.8–6.0
Bromothymol blue	6.0–7.6
Phenolphthalein	8.2–10.0
Alizarin yellow	10.2–12.0

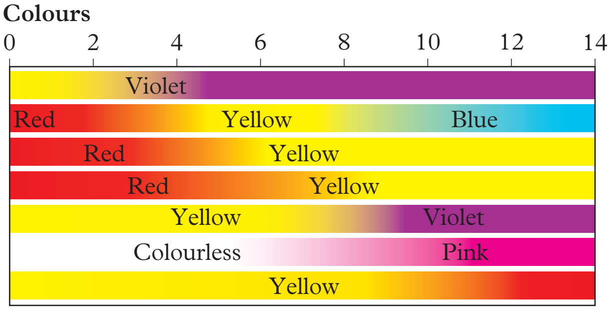


FIGURE 15 Acid–base indicators change colours at different pH values.

Reactions with metals

You may recall from Year 11 Chemistry that acids react with metals to produce an ionic salt and hydrogen gas. The pop test, where magnesium ribbon reacts with hydrochloric acid in a test tube, is a common example of this. An inverted test tube is placed on the top of the reaction tube to collect the gas that is generated in the reaction. To verify that the gas is hydrogen, a highly flammable and explosive gas, a match is used to ignite the gas in the inverted test tube, resulting in a pop sound when the gas ignites.

$$\mathrm{Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)}$$

The presence of a flammable gas or the increase in the pH of the reaction mixture, due to a decrease in H^+ (from HCl(aq)) when H_2 gas is formed, can be used to test for acids. The same reaction is possible for organic acids.

Figure 16 demonstrates the reaction between magnesium ribbon and aqueous solutions of carboxylic acids. The hydrogen ion in ethanoic acid is displaced by the magnesium forming hydrogen gas. As these are weak acids, they do not completely dissociate, resulting in the following reaction.

$$\mathrm{2CH_3COOH(aq) + 2Mg(s) \rightleftharpoons 2CH_3COO^-(aq) + 2Mg(s) + H_2(g)}$$

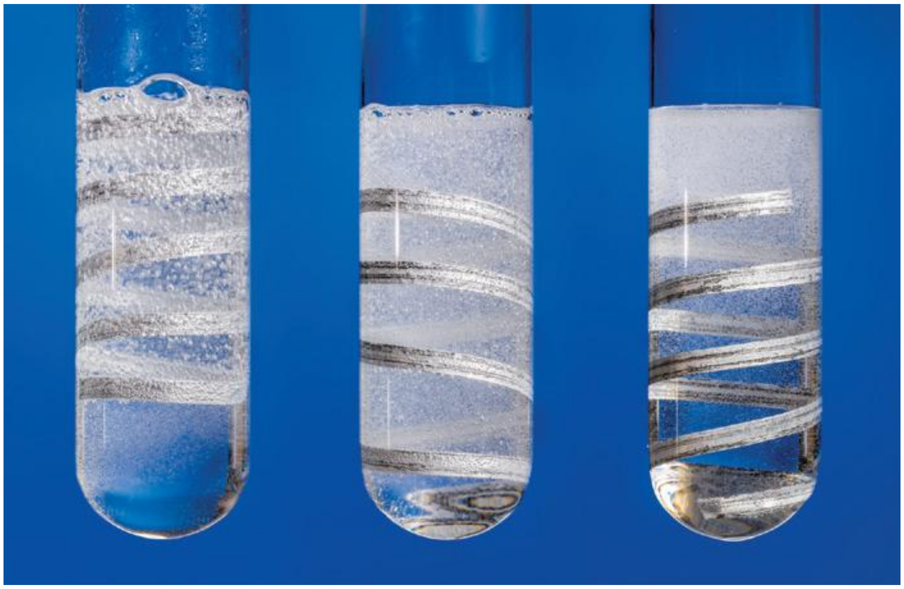


FIGURE 16 The reactions between magnesium ribbon and aqueous solutions of methanoic acid (left), ethanoic acid (middle) and propanoic acid (right).

TABLE 1 Common esters and their fragrances

Ester	Scent or flavour
Benzyl acetate	Jasmine
Butyl acetate	Raspberry
Ethyl butanoate	Pineapple
Octyl acetate	Orange
Pentyl acetate	Banana
Ethyl salicylate	Mint
Ethyl cinnamate	Cinnamon

Esterification reactions

In Chapter 9, you learnt that carboxylic acids react with alcohols to form esters with a sulfuric acid catalyst. Table 1 outlines the common fragrances produced by esters.

Reacting an alcohol and a sulfuric acid catalyst may not be able to identify the specific carboxylic acid, but it can be used to identify the presence of a carboxylic acid if an ester is formed.

10.1 SKILL DRILL

Analysing qualitative data to identify unknown organic compounds

Key science skill: Analyse and evaluate data and investigation methods

A laboratory scientist starts a new job and enters the laboratory that she will be working in. Unfortunately, the last scientist has left four unlabelled containers of chemicals on the bench. She knows that she cannot simply dispose of the chemicals down the sink as they may be toxic, corrosive, carcinogens or harmful to aquatic life. Instead, she runs a series of basic chemical tests using small samples of each chemical to see if she can identify them quickly and efficiently, before more complex methods are required. She runs the following tests and obtains the results in the table below.

Test	Summary
1	Bromine is added to the sample in a test tube.
2	Permanganate is added to the sample in a tube.
3	Hydrochloric acid is added.
4	Sodium hydroxide is added.
5	Ethanol with sulfuric acid is added.

Results

Sample	Chemical test				
	1	2	3	4	5
A	Stays red	Turns colourless	No reaction	No reaction	No reaction
B	Stays red	Stays purple	No reaction	Bubbles form	Fruity smell
C	Turns colourless	Turns colourless	No reaction	No reaction	No reaction
D	Turns colourless	Stays purple	No reaction	Bubbles form	Fruity smell

Practise your skills

1 Explain what you can conclude about the samples A–D.

2 Contrast qualitative and quantitative data.

3 Describe the benefits of each type of data in the context of the question.

Need help analysing and evaluating data? See Topic 1.8 (page 24).

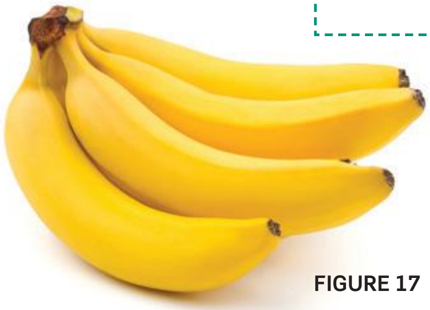


FIGURE 17 The ester pentyl acetate smells like banana.

See how an understanding of the different qualitative tests can be used to select the best methods to identify unknown compounds in Worked example 10.1.

10.1 CHALLENGE

Identifying multiple functional groups

A chemist is asked to identify the organic molecule in a solution that is labelled as 'X'. The chemist suspects that the molecule may be lactic acid (Figure 18).

Describe the qualitative tests that could be performed on this molecule and what results you would expect to see.

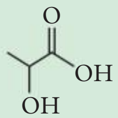


FIGURE 18 Lactic acid

10.1 Worked example
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10.1 Worked example
Video demonstration

10.1 CHECK YOUR LEARNING

Describe and explain

1 Using balanced chemical equations, explain the tests that can be used to determine whether a molecule has carbon–carbon double bonds.

2 Using examples, explain the difference between a qualitative and a quantitative test.

Apply, analyse and compare

3 Using balanced chemical equations, compare the oxidising agents used to test for alcohols and their respective colour changes.

4 A chemist must conduct a series of analyses to confirm the identity of a chemical as but-2-ene.

  a Using balanced chemical equations, explain the tests that would confirm the identity of the molecule.

  b Explain what cannot be confirmed using these experimental techniques.

5 A chemist must conduct a series of analyses to confirm the identity of a chemical as 2-hydroxypentanoic acid.

  a Using balanced chemical equations, explain what tests would confirm the identity of the molecule.

  b Explain what cannot be confirmed using these experimental techniques.

Design and discuss

6 Design a series of experiments and place them into a flowchart to determine the identity of a substance you suspect is ethanoic acid. For each step, state what you are testing for and how you would know if the test is positive.

FIGURE 19 Many qualitative tests conducted in the laboratory rely on observations (like the colour of the solution) to characterise compounds.

10.2 Purity of consumer products

KEY IDEAS

In this topic, you will learn that:

- melting and boiling points can be used to verify the identity and purity of solids
- distillation separates components of mixtures based on their boiling point.

Before they make it onto the market shelves, consumer products must be tested to verify their identity and purity. This is especially important for food products or other chemicals that are used on the body. The presence of potentially harmful contaminants or other unwanted compounds can be detrimental to the function or safety of the product.

There are a number of ways in which identity and purity can be determined. Previously, you learnt about some qualitative analysis methods. In this topic, you will learn about how the melting and boiling points of substances can be used to verify the identity and purity of a substance.

FIGURE 1 Iron has a melting point of 1538°C.

Factors affecting melting and boiling points

Intermolecular bonding and the interactions between molecules determine the physical properties and behaviours of all molecules, including melting point and boiling point. Molecules that have stronger intermolecular forces have higher melting and boiling points due to the large amounts of energy required to change their state. This can be explained using the kinetic theory of matter.

kinetic energy the energy involved with the motion of particles, which can be increased by adding heat energy

All particles have **kinetic energy**. The greater the kinetic energy of a particle, the faster it moves. When investigating the melting and boiling points of a substance, it is important to consider:

- the kinetic energy of the particle: the more kinetic energy a particle has, the more it moves around and breaks away from other particles
- the energy involved with the intermolecular forces holding the molecules together: the greater the intermolecular forces holding the particles together, the harder it is to break them apart. If the energy of particle movement is greater than the energy of the attraction, the particle will break free.

Melting point

The melting point of a substance is the temperature at which it changes from solid into liquid state. In solid state, intermolecular forces between molecules are so great that they are held together in a rigid structure. Their motion is so limited that they only vibrate in place.

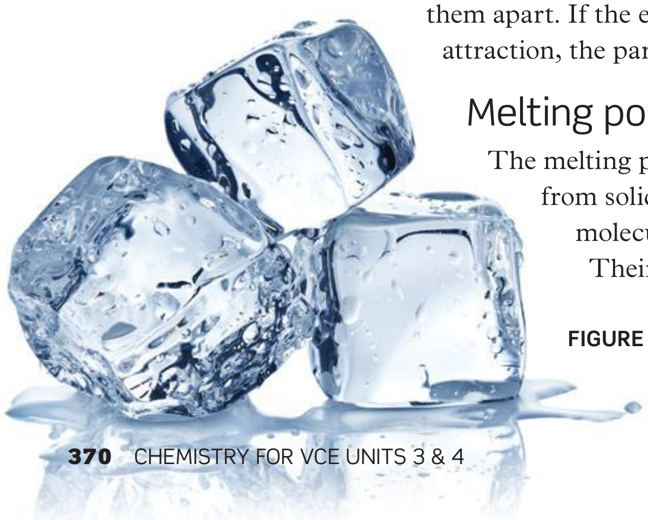


FIGURE 2 At the melting point of ice, these cubes turn to liquid.

When heat energy is applied, the molecules of a solid begin to move faster, and eventually gain enough energy to start moving away from neighbouring molecules. The molecules are still attracted to one another but are not longer held together as tightly. The temperature at which heat energy is sufficient to allow molecules to break free of their rigid structure and become liquid is the melting point.

Boiling point

The boiling point of a substance is the temperature at which it changes from liquid into gaseous state. In liquid state, intermolecular forces between molecules are strong enough to hold molecules together and prevent them from fully breaking apart from each other. The molecules move over and around each other, but the kinetic energy is not great enough for them to break free.

When heat energy is applied, molecules begin to move faster and eventually gain enough kinetic energy to start breaking free from neighbouring molecules. The molecules are no longer held to each other because their kinetic energy is greater than the energy involved with their intermolecular bonding (Figure 3). The temperature at which heat energy is sufficient to allow molecules to break free from each other into a gaseous state is the boiling point.

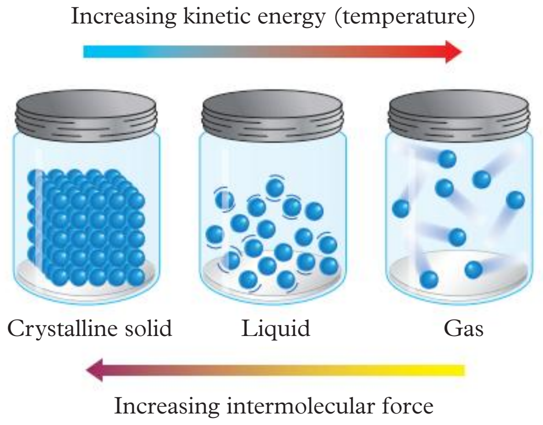
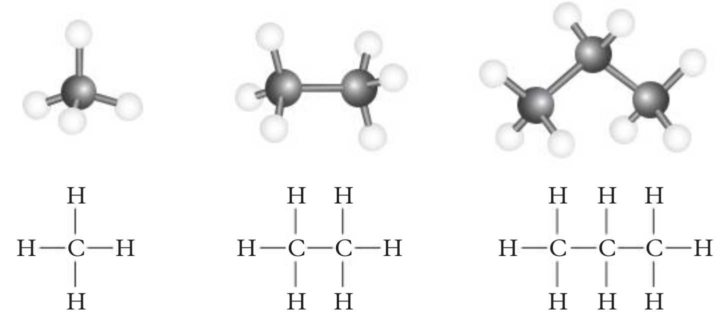


FIGURE 3 The relationship between kinetic energy and intermolecular bonding in solid, liquid and gaseous states

Trends in melting and boiling points

The stronger the intermolecular bonding between molecules (dispersion forces, dipole–dipole interactions and hydrogen bonding), the more kinetic energy and, therefore, heat is required to break the molecules apart. If molecules only interact through weak dispersion forces, then less kinetic energy and therefore heat is required to break them apart. Let's recap the trends in melting and boiling points.

The carbon atoms in alkanes are covalently bonded in a long chain of single bonds, and the remaining bonds are to hydrogen.



Methane (CH_4) **Ethane (C_2H_6)** **Propane (C_3H_8)**

FIGURE 4 Structures of the common alkanes: methane, ethane and propane

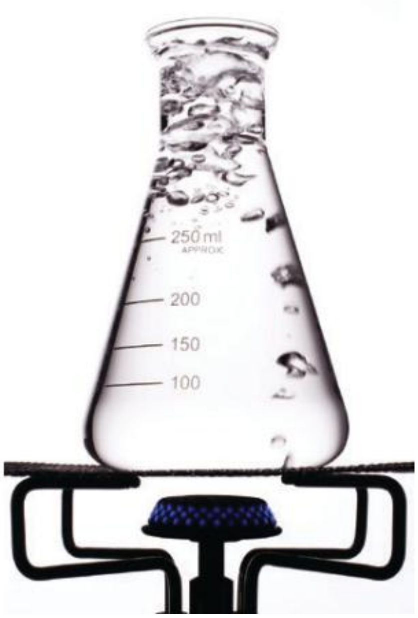


FIGURE 5 Substances with weaker intermolecular forces have a lower boiling point.

Alkanes experience only dispersion forces because they are non-polar. Figure 6 shows the relationship between the number of carbon atoms in an alkane and its boiling point. As the number of carbon atoms increases, the number of electrons increase and the dispersion forces increase. As more heat energy is needed to separate the weakly interacting molecules, the boiling point of the molecule also increases.

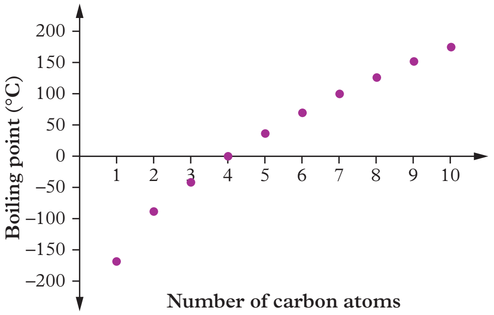
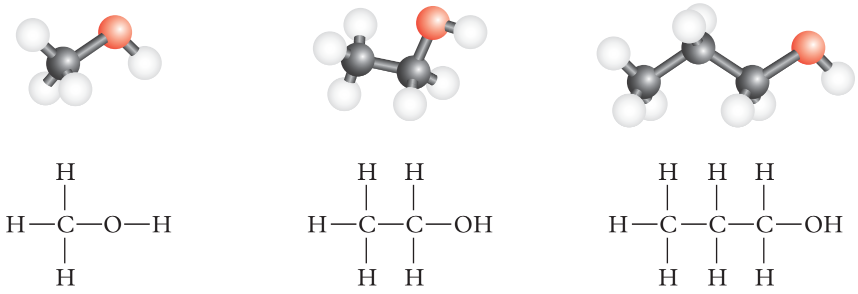


FIGURE 6 Boiling points of straight-chain alkanes containing 1–10 carbon atoms

What about primary alcohols? The presence of the electronegative oxygen atom and the shape of the molecule make primary alcohols polar, allowing them to undergo hydrogen bonding with other alcohol molecules. However, because primary alcohols contain a carbon chain, one end of the molecule experiences a small amount of dispersion forces.



Methanol (CH_3OH) **Ethanol (C_2H_5OH)** **1-propanol (C_3H_7OH)**

FIGURE 7 Structures of the primary alcohols: methanol, ethanol and 1-propanol

All primary alcohols have the same amount of hydrogen bonding because they have the same polar –OH group. Therefore, any differences in boiling points are not due to the polarity of the –OH group, but rather to the length of the carbon chain and therefore the amount of dispersion forces (Figure 8).

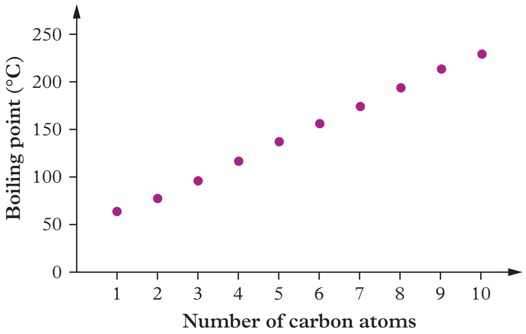


FIGURE 8 Boiling points of primary alcohols containing 1–10 carbon atoms

When the boiling points for alkanes are plotted together with the boiling points of primary alcohol, the trend between molecules that have different intermolecular forces can be seen (Figure 9). Polar molecules that experience hydrogen bonding are more attracted to one another. They are harder to break apart than molecules that contain dispersion forces.

Converting molecules with only dispersion forces (e.g. alkanes) from a liquid to a gaseous state requires less energy than it does for primary alcohols. Therefore, the boiling point of alkanes is lower than that of primary alcohols. The melting point of covalent molecules is similarly based on the strength of the intermolecular forces between molecules. Worked example 10.2 walks you through how to compare the melting points of two different substances.

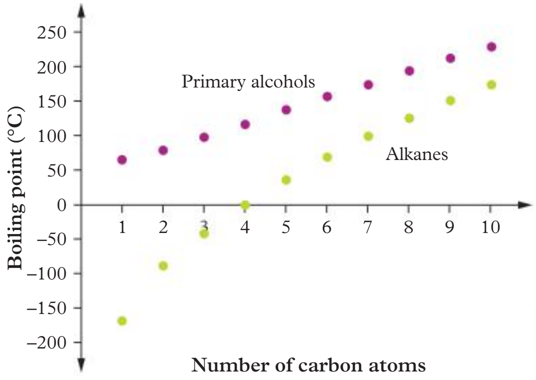


FIGURE 9 Boiling points of straight-chain alkanes, containing dispersion forces only, and primary alcohols, containing hydrogen bonds, and dipole–dipole and dispersion forces

Study tip

Plotting data on the same axis allows you to do more a complex analysis. Think about how you can bring together different data to enrich your discussion.

10.2 Worked example
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10.2 Worked example
Video demonstration

chemical purity
a measurement of the absence of impurities in a sample

theoretical value
the values that have been determined using specialised equipment and chemicals of the highest purity; also known as a literature value

Determining chemical purity

Chemical purity, the absence of impurities in a sample, mixture or solution, can be determined using melting and boiling points, as these values are unique to each compound. By experimentally determining the melting and boiling points, we can verify the identity of a compound against **theoretical values**, as well as determine the purity of the compound.

Using melting point to determine purity

Pure compounds have sharp melting points (narrow range). On the other hand, impure substances generally melt within a larger temperature range and lower the melting range of a chemical compound. This is because they disrupt the solid crystalline structure that holds a solid chemical substance together. This lowers the amount of energy and, therefore, heat required to disrupt the intermolecular forces between molecules or ions.

For example, alloyed metals tend to have lower melting points than pure metals. The 9ct yellow gold used by jewellers is an alloy that consists of 37.5% pure gold, with the remaining percentage a mixture of other metals such as zinc, copper, iron and silver. The melting point of 9ct gold falls between 970°C and 1040°C, whereas pure gold has a melting point of 1945°C.

Melting points can be measured using a melting point apparatus (Figure 10). It uses a capillary tube filled with a small sample of solid. The temperature of the solid is gradually increased until it melts. There is a magnifying window in the apparatus to view the bottom of the capillary tube. This allows you to see when the solid substance first begins to melt and when it has fully melted. The temperature recorded is typically a range.

FIGURE 10 A melting point apparatus

The method for preparing a solid for the melting point apparatus is shown in Figure 11 and the stages of melting a solid are shown in Figure 12.

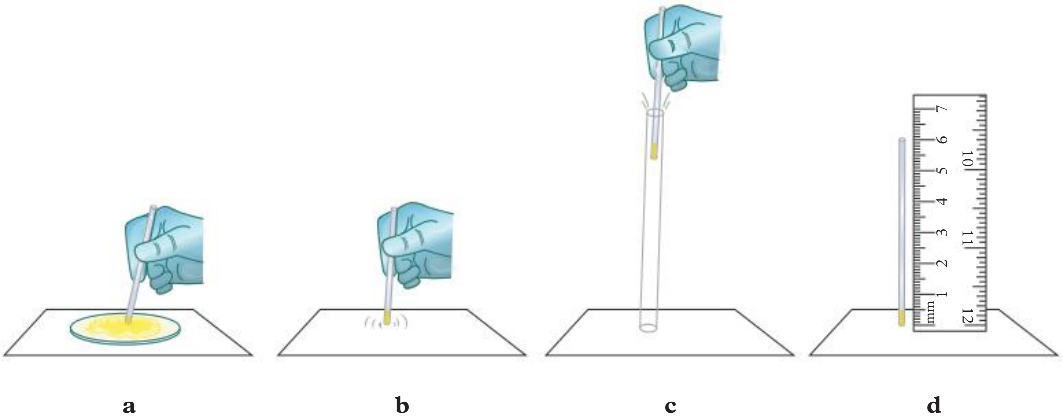

FIGURE 11 **a** The solid sample is collected in the capillary. **b** It is inverted and tapped to allow the solid to move to the bottom of the tube. **c** The tube is dropped through a longer tube, using gravity to help the solid pack tightly at the bottom. **d** The amount (height) of the sample in the tube is measured using a ruler.

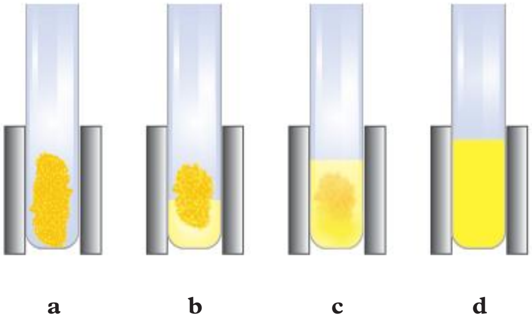

FIGURE 12 The melting of benzoic acid. **a** The crystals are visible at the bottom of the capillary tube. **b** The temperature at which the first visible drop is seen is recorded. **c** You need to carefully observe the sample and note down the final temperature when **d** the whole sample has melted.

Study tip

A higher melting point is caused by stronger intermolecular forces (stronger attractions between molecules).

Study tip

Substances will vaporise at lower temperatures if they have weaker intermolecular forces.

The closeness of the experimental melting point to the theoretical value gives an indication of how pure the sample is. If the value is slightly lower, or over a larger range, it is likely there are impurities present. And if the melting point is way off? You might have a completely different compound!

Using boiling point and distillation to determine purity

distillation
a separation technique where components of a mixture are vaporised, rise and pass through a condenser where they are turned back into a liquid, and are then collected, having been separated from the mixture

Distillation is the separation of a mixture using the boiling points of its components. As with the melting point, pure substances should have a clear boiling point. You would therefore expect impure substances to have a wider boiling point range.

The simplified steps of distillation are as follows:

1. A liquid mixture of compounds is heated in a sealed container.
2. The components of the mixture will vaporise and form a gas in order of increasing boiling point: the component with the lowest boiling point will vaporise first.
3. When a component forms a gas, it is captured and condensed back into a liquid, which can be collected.

There are two types of distillation: simple and fractional.

Simple distillation

Simple distillation is used to separate mixtures with components that have significantly different boiling points. A mixture of substances is placed in the round bottom flask and heated using a heating element such as a Bunsen burner. More advanced apparatuses use electric heating elements that can heat and maintain the mixture at specific temperatures. The components of the mixture that vaporise rise from the distilling flask into the condenser. The condenser is cooled by a constant flow of cool water and causes the vapour captured to condense back into a liquid and drip down into the conical flask. This is summarised in Figure 13.

Fractional distillation

Fractional distillation is used to separate mixtures with components that have relatively similar boiling points into different fractions. The mixture is first heated to the lowest specific temperature and the fraction of the mixture that vaporises at this temperature is captured as it exits the condenser. The temperature of the heating element is increased, and a new fraction is captured as it exits the condenser. The process is repeated until all fractions have been separated according to their temperature.

Fractional distillation is widely used in the purification of crude oil, a fossil fuel containing a mixture of hydrocarbon atoms which can be separated and used for a variety of different purposes (Figure 14). The crude oil mixture is passed through a furnace to be boiled at over 600°C. The boiling oil is then passed into the bottom of the fractional distillation column. As soon as the mixture enters the column, more volatile substances (those that vaporise at lower temperatures) will rise and move into higher compartments. The smaller the molecule and the weaker its intermolecular forces, the lower its boiling point and the higher it rises in the fractional distillation column.

simple distillation
a separation technique where components of a mixture are vaporised, condensed as they are cooled by running water and collected in a distilling flask

fractional distillation
a separation technique where components of a mixture are vaporised, rise through a column until they reach a compartment of a temperature lower than their boiling point, condense and are removed from the column

Study tip

Use simple distillation for compounds with very different boiling points. If they are very similar, use fractional distillation.

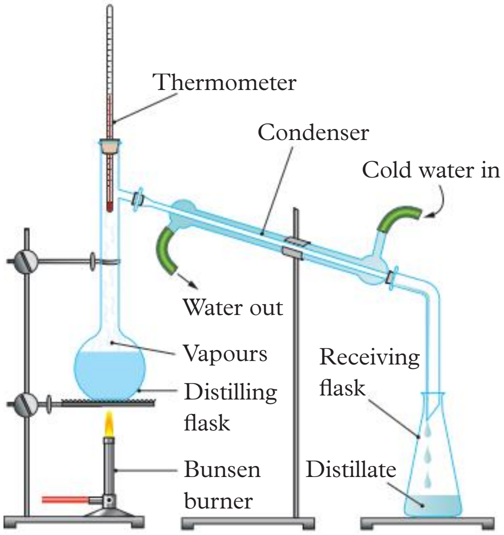

FIGURE 13 A distillation apparatus

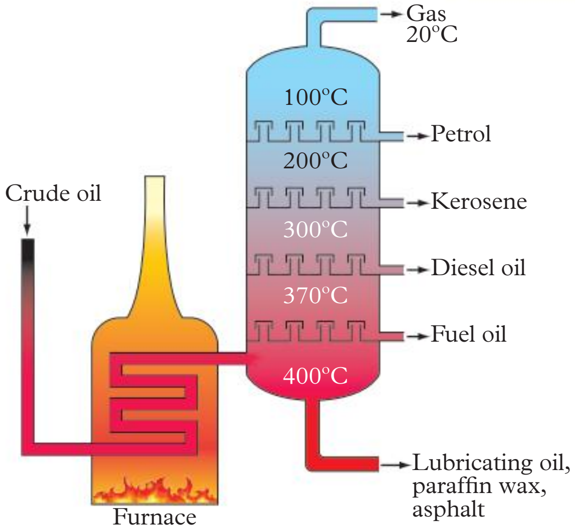

FIGURE 14 Fractional distillation is used to purify crude oil.

As the crude oil vapour moves up the column, the environment becomes cooler. Vaporised molecules will rise in the column until they move into a compartment where the temperature is lower than its boiling point. Once this happens, the vapour will condense and be captured in the tray of that compartment.

Explore the melting and boiling points of fats in Real-world chemistry 10.2.

10.2 Real-world chemistry
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10.2 SKILL DRILL

Analysing the boiling points obtained from fractional distillation

Key science skill: Analysing and evaluating data and investigation methods

A mixture of primary alcohols from C1 (containing one carbon atom) to C5 (containing five carbon atoms) is separated using fractional distillation. The theoretical boiling points of the alcohols are:

TABLE 1 Boiling points of primary alcohols

Number of carbon atoms	Boiling point (°C)
1	64.7
2	78.4
3	97.0
4	117.7
5	138.0

The student uses a Bunsen burner as the heating source in the distillation and separates the mixture into two fractions. She concludes that one fraction was maintained below 100°C and the other was above 100°C.

Practise your skills

1 Comment on whether this analysis is qualitative or quantitative and explain why.

2 State an error in the experiment and explain its impact on the experiment outcome.

3 Explain how this error can be minimised, referring specifically to how it would improve the outcome of the experiment.

Need help analysing and evaluating investigation methods? See Topic 1.8 (page 24).

10.2 CHECK YOUR LEARNING

Describe and explain

1 Explain why kinetic energy affects the ability of a substance to melt or boil.

2 Explain the difference between melting point and boiling point by using the terms *kinetic energy* and *intermolecular forces*.

3 Describe how melting points differ between pure and impure chemical substances.

Apply, analyse and compare

4 Compare fractional distillation with simple distillation.

5 A mixture of the following substances was separated using fractional distillation: $CH_3CH_2OH(l)$, $CH_3OH(l)$, $CH_3CH_2CH_2OH(l)$. Determine the order in which they would form fractions, from the lowest temperature fraction to the highest.

6 A mixture of the following substances was separated using fractional distillation: $CH_3CH_3(l)$, $CH_3CH_2OH(l)$, $CH_3CH_2COOH(l)$. Determine the order that they would form fractions from the highest temperature fraction to the lowest.

Design and discuss

7 Sugar is made of sucrose, a disaccharide of fructose and glucose molecules bonded together. In the lolly-making process, sugar starts to melt at 160°C and caramelises at 170°C.

  a Discuss why sugar has such a high melting point.

  b Discuss what happens to the intermolecular forces between sucrose molecules and to the structure of sucrose as it melts.

8 Discuss the importance of melting and boiling points in verifying the components and purity of consumer products. Research one consumer product where this type of analysis would be useful and use as an example in your answer.

9 Evaluate the following claim: *Melting point determination could be considered both a qualitative and quantitative technique*. Identify whether you think the claim is correct or incorrect and justify your response.

10.3 Iodine titrations to test unsaturation

KEY IDEAS

In this topic, you will learn that:

- volumetric analysis is used to calculate unknown concentrations and volumes
- the degree of unsaturation can be identified through an iodine test or titration.

In Year 11, you learnt about volumetric analysis, including acid–base titrations. Titrations can also be used to analyse organic compounds. In this topic, you will learn how to use the results of iodine titrations to determine the degree of unsaturation. Then, in Topic 10.4, you will learn about redox titrations for analysing organic compounds that can undergo oxidation or reduction. But first, let's recap the basics.

Volumetric analysis

volumetric analysis
a quantitative analytical technique where the volume and concentration of a solution are accurately known and used to determine the concentration of an analyte which also has an accurately known volume

titration
the addition of a standard solution to a solution of unknown concentration until the reactants have reacted in equal mole amounts

standard solution
a solution with a highly accurately known concentration

Volumetric analysis is an essential chemical technique used to determine the concentration or amount of a solute in a larger sample. It measures highly accurate volumes of solutions and produces results with high accuracy and reliability. Volumetric analysis is used in food chemistry to determine amounts of different nutrients, the analysis of minerals in mining, the analysis of nutrients in seawater and many more.

This type of analysis uses stoichiometry to determine the concentration of an unknown solution by titrating it against a solution of known concentration. **Titrations** are based on the same general principle, that a **standard solution** of highly accurate concentration can be used to determine the concentration of a solution of unknown concentration using stoichiometry.

FIGURE 1 Volumetric analysis uses stoichiometry to find the concentration of an unknown solution by titrating it against a known solution.

Primary standards and standard solutions

volumetric flask the piece of glassware used to create a standard solution due to its ability to measure highly accurate volumes

primary standard the solute that is dissolved in a solvent to create a standard solution

To perform a titration, you need to prepare a standard solution. Remember that standard solutions are prepared in a **volumetric flask**. A solute is selected as the **primary standard**, which is then weighed and dissolved in a solvent. A primary standard has several important qualities; it must:

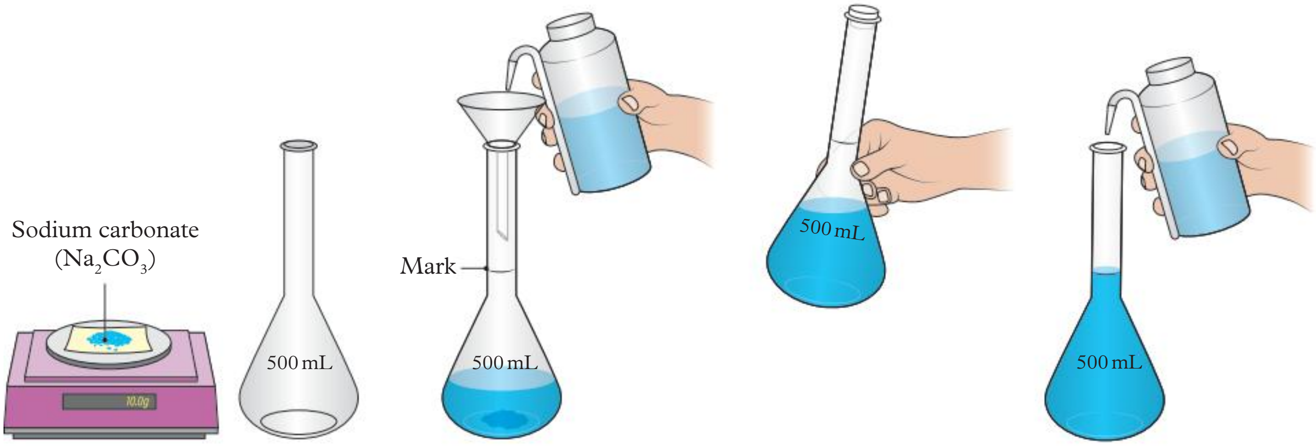
- be a solid that is soluble in water, so it can be dissolved to make a standard solution
- be highly pure
- have a known chemical formula
- have a high molar mass, so errors during weighing are minimised
- be stable or unreactive so as not to absorb water, carbon dioxide or other chemicals from the atmosphere or surrounding environment
- be anhydrous. For example, $MgSO_4$•$7H_2O$ (hydrated magnesium sulfate or epsom salt) is not an appropriate standard because the seven water molecules can evaporate, changing the chemical formula and molar mass
- be reasonably priced.

FIGURE 2 A volumetric flask

Preparing a standard solution

The following steps are used to prepare a standard solution:

1. Rinse the volumetric flask with water and then with the desired solvent.
2. Weigh the primary standard on an electronic balance and record its mass.
3. Add the primary standard to a volumetric flask, using a filter funnel.
4. Half-fill the volumetric flask with solvent, put the lid on the flask and swirl the flask until the standard is completely dissolved.
5. Use a wash bottle to make the solution up to the mark on the flask.
6. Place the lid back on the flask and hold it firmly in place while you tip the flask upside down and then upright again at least five times, ensuring that the solution has an evenly distributed concentration.



a Weigh out the primary standard on an analytical balance.

b Add the standard to the volumetric flask and then half-fill the flask with solvent.

c Put the lid on the flask and swirl it to dissolve the solid.

d Add more solvent up to the mark.

FIGURE 3 Preparing a standard solution

The formulas below can be used to calculate the molar concentration of the standard solution, where n is the amount of solute in a solution (mol) and c is the concentration of the solution (M).

$$n = \frac{\text{mass of solute (g)}}{\text{molar mass (g mol}^{-1}\text{)}} = \frac{m}{M}$$

$$c = \frac{\text{amount of solute (mol)}}{\text{volume of solution (L)}} = \frac{n}{V} = \frac{m}{M \times V}$$

Worked examples 10.3A and 10.3B demonstrate the calculations required for preparing standard solutions. If you're feeling confident, try solving a trickier problem in Challenge 10.3.

10.3 Challenge
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10.3A WORKED EXAMPLE

CALCULATING THE MASS REQUIRED TO MAKE A STANDARD SOLUTION

Students are instructed to make a 0.20 M standard solution of aqueous oxalic acid ($H_2C_2O_4$) in a 250 mL volumetric flask. Calculate the mass of oxalic acid they should use.

Solution

Think	Do
Step 1: Using concentration and volume, calculate the amount, in mol, of oxalic acid.	$n(H_2C_2O_4) = c \times V$ $= 0.20 \times (250 \div 1000)$ $= 0.05\,\text{mol}$
Step 2: Using the amount and molar mass, calculate the mass of oxalic acid.	$m(H_2C_2O_4) = n \times M$ $= 0.05 \times ((2 \times 1.0) + (2 \times 12.0) + (4 \times 16.0))$ $= 4.5\,\text{g}$ (2 sig fig)

10.3B WORKED EXAMPLE

CALCULATING THE CONCENTRATION OF A STANDARD SOLUTION

A standard solution is made from 1.25 g of potassium hydrogen iodate (HI_2KO_6) in water in a 20.00 mL volumetric flask. What is the concentration of the solution?

Solution

Think	Do
Step 1: Calculate the amount, in mol, of potassium hydrogen iodate, using mass and molar mass.	$n(HI_2KO_6) = \frac{m}{M}$ $= \frac{1.25}{1.0 + (2 \times 126.9) + 39.1 + (6 \times 16.0)}$ $= 3.2 \times 10^{-3}\,\text{mol}$
Step 2: Use the amount, in mol, and volume to calculate the concentration of potassium hydrogen iodate. To get a precise value, avoid rounding the number of moles from Step 1.	$c(HI_2KO_6) = \frac{n}{V}$ $= \frac{3.2 \times 10^{-3}}{20.00 \div 1000}$ $= 0.16\,\text{M}$ (2 sig fig)

burette
graduated glass cylinder that dispenses high-accuracy volumes of standard solution

pipette
a glass tube that dispenses high-accuracy volumes of a solution with unknown concentration

aliquot
the volume of solution dispensed, using a pipette, into the conical flask

indicator
a chemical substance that can be chosen to change colour when the equivalence point of a reaction is reached

equivalence point
the point in a titration when the chemicals have reacted in the molar ratio of the balanced chemical equation

titrant
the name given to the solution in the burette

end point
the change in colour of the indicator during a titration

titre
the volume of solution dispensed from the burette

concordant titres
titres within 0.1 mL from the highest to lowest volume

Study tip

Concordant titres must be 0.1 mL between the highest and lowest values. This ensures precision between repeats.

10.3
Skill drill
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Performing a titration

To perform a titration, you need specialised equipment – a **burette** and a **pipette**. Burettes and pipettes dispense highly accurate volumes of solution, increasing the accuracy of the analytical technique. The general method to perform a titration is:

1 Use a pipette to transfer a volume (**aliquot**) of the solution being analysed (analyte) into a conical flask. At this point, an **indicator** can be added into the conical flask, which shows a colour change at the **equivalence point** of the reaction.
2 Fill the burette with the standard solution, making sure that there are no bubbles. This solution is called the **titrant**. Record the initial volume of the burette.
3 Turn on the burette to start dispensing the titrant into the solution in the conical flask. This should start to cause a colour change, which disappears with stirring. Stir the conical flask constantly to ensure that the reaction is evenly mixed throughout the flask.
4 Continue to add the titrant until the colour becomes harder to remove with stirring. At this point, turn the tap so that it only dispenses one drop at a time.
5 Turn off the tap when a permanent colour change is observed. This is the **end point**.
6 Record the final volume in the burette. Subtract the initial volume from the final volume to determine the total volume of solution dispensed from the burette. This is the **titre**.
7 Repeat the method until you have obtained three **concordant titres**. Concordant titres differ by 0.1 mL from lowest to highest; 1 drop of solution has a volume of 0.05 mL, so all titres should be within two drops of each other. The average of the concordant titres is used in stoichiometric calculations.

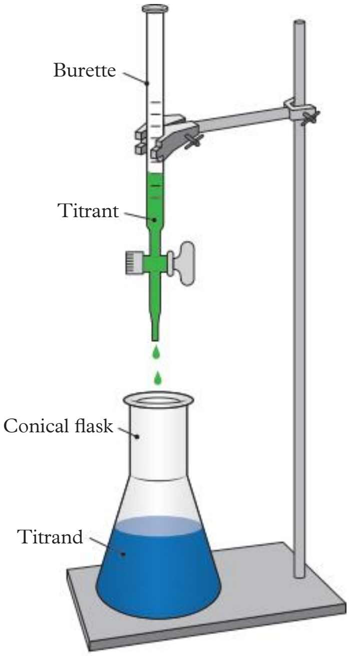

FIGURE 4 Titration apparatus

You might remember from Year 11 that there are many potential errors when conducting titrations. Apply your evaluation skills to identify errors in Skill drill 10.3. If you're not feeling confident, review your understanding of Chapter 15 in the *Chemistry for VCE Units 1 and 2 Student Book*.

Testing for the degree of unsaturation with iodine

In Topic 10.2, you learnt about how the bromine test can be used to identify carbon–carbon double bonds in organic compounds. Saturated and unsaturated compounds can also be distinguished using iodine testing. Iodine testing can qualitatively determine the presence of double bonds, but it can also be considered a **quantitative test** that provides information about the degree of unsaturation.

quantitative test
a test that gathers measurable data, expressed numerically

The iodine test

The **iodine test** is based on a similar principle to the bromine test, where iodine molecules (I_2) will readily break carbon–carbon double bonds and attach to the carbon atoms at the site of these bonds (Figure 6). In the presence of carbon–carbon double bonds, the dark brown iodine will react and become colourless. Depending on the amount of iodine reacted, and therefore the amount of double bonds in the organic compound, the colour intensity of the solution will differ. This provides qualitative data, as we observe the change in colour of the solution.

iodine test
a qualitative test in which iodine is added to a hydrocarbon sample to test for carbon–carbon double bonds; if the iodine changes from brown to colourless, then the hydrocarbon is unsaturated

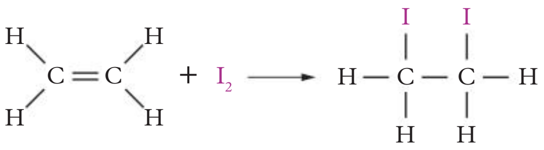

FIGURE 6 Iodine molecules break carbon–carbon double bonds and attach at the site of these bonds.

FIGURE 5 Iodine will react with unsaturated organic compounds.

Titrations to determine iodine value

Alternatively, the sample being analysed can be titrated against iodine, with a starch indicator. This forms an **iodohydrocarbon**. A known concentration and volume of analyte is placed in the conical flask with a starch indicator. Iodine, with a known concentration, is added until a colour change is observed. The volume is recorded and the amount of iodine reacted can be calculated.

iodohydrocarbon
an organic compound consisting of carbon, hydrogen and iodine atoms

The more carbon–carbon double bonds within a molecule, the more iodine must be added to reach the equivalence point and the greater the degree of unsaturation. As one I_2 molecule will add across one double bond, the molar ratio of the balanced chemical equation will be altered in the following pattern:

$$1 \text{ double bond} + I_2 \rightarrow \text{iodohydrocarbon}$$
$$2 \text{ double bonds} + 2I_2 \rightarrow \text{iodohydrocarbon}$$
$$3 \text{ double bonds} + 3I_2 \rightarrow \text{iodohydrocarbon}$$

The number of double bonds (or the degree of unsaturation) can be determined based on the stoichiometric ratio of iodine to hydrocarbon:

$$\text{Degree of unsaturation} = \frac{n(I_2)}{n(\text{hydrocarbon})}$$

This is considered quantitative data, as the number of double bonds present in the compound is measured and a numerical value is obtained. This can then inform the identification of the organic compound. Worked examples 10.3C and 10.3D demonstrate how to calculate the degree of unsaturation following titration against iodine.

Study tip

The method for the iodine titration is not within the scope of the VCE Chemistry Study Design, so you do not need to know the details. What you do need to know is how to use stoichiometry to calculate the degree of unsaturation.

10.3C Worked example
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10.3C Worked example
Video demonstration

10.3D WORKED EXAMPLE

CALCULATING THE DEGREE OF UNSATURATION FROM THE VOLUME OF IODINE REACTED

0.010 mol of a fatty acid, an unsaturated carboxylic acid, reacts with exactly 10.0 mL of 5.00 M I_2. Determine the number of double bonds in the fatty acid.

Solution

Think	Do
Step 1: Calculate the amount, in mol, of iodine, I_2.	$n(I_2) = c \times V$ $= 5.00 \times (10.0 \div 1000)$ $= 0.050$ mol
Step 2: Divide the amount in mol of I_2 by the amount in mol of hydrocarbon to calculate the degree of unsaturation.	Degree of unsaturation $= \frac{n(I_2)}{n(\text{hydrocarbon})}$ $= \frac{0.050}{0.010}$ $= 5$ This means that the hydrocarbon contains five carbon–carbon double bonds.

10.3 CHECK YOUR LEARNING

Describe and explain

1 Describe the difference between a primary standard and a standard solution.

2 Describe the difference between an end point and an equivalence point.

3 Explain what the degree of unsaturation is.

4 Name the type of reaction that occurs when iodine is reacted with an unsaturated hydrocarbon.

Apply, analyse and compare

5 Calculate, to the correct number of significant figures:

- **a** the concentration of a standard solution when 0.8 g of $KMnO_4$ is dissolved in 100 mL of water in a volumetric flask
- **b** the mass required to produce an aqueous 0.50 M standard solution of sodium oxalate ($Na_2C_2O_4$) in a 250 mL volumetric flask.

6 Calculate the degree of unsaturation of a hydrocarbon when 0.05 mol of the hydrocarbon react with exactly 100.0 mL of 2.00 M I_2.

7 An unsaturated hydrocarbon is hypothesised to contain three double bonds. If 5.00 mol of the hydrocarbon are reacted with iodine that has a 1.5 M concentration, calculate the volume of iodine that will react.

Design and discuss

8 A student performs a titration and forgets to rinse the burette with standard solution before filling it. Predict the effect on the concentration of the analyte, using a calculation to support your answer.

9 When weighing a primary standard on an electronic balance, a student neglects to tare the balance. Before adding the solid to the weigh boat, the balance reads –0.50 g. Determine the effect of this error on:

- **a** the concentration of the standard
- **b** the amount, in mol, of the standard that will react with the analyte
- **c** the concentration of the analyte.

10.4

Redox titrations for volumetric analysis

KEY IDEAS

In this topic, you will learn that:

- redox titrations can be used to quantitatively analyse organic compounds
- addition of an indicator solution is not required for self-indicating redox reactions.

Many of the ingredients in consumer products are organic compounds. For example, ethanol and sulfur dioxide are present in wine; vitamin C or ascorbic acid is found in orange juice; hydrogen peroxide is the antiseptic in mouthwash and other cleaning products. Depending on their structure, they may be able to undergo oxidation or reduction.

Alcohols can be oxidised into aldehydes, ketones and carboxylic acids. The reverse is also true: carboxylic acids, ketones and aldehydes can be reduced back into alcohols. Their ability to undergo redox reactions means that they can be analysed by redox titrations.

Redox titrations

redox titration
titrations that involve redox reactions

Redox titrations follow the same general principles as acid–base titrations, but analyse redox reactions. This type of analysis is quantitative, as the amount of organic compound in the sample can be determined.

The equivalence point is a measure of a stoichiometric ratio

The equivalence point of a titration is reached when chemicals have reacted according to their molar ratio. This means that the limiting reactant has been completely consumed and what remains in the conical flask are the products, and any excess reactant. Consider the reaction between tin(II) ions and iodine:

$$Sn^{2+}(aq) + I_2(aq) \rightarrow Sn^{4+}(aq) + 2I^-(aq)$$

One mole of Sn^{2+} will react with one mole of I_2 to produce one mole of Sn^{4+} and two moles of I^-. If the amounts of Sn^{2+} and I_2 before the reaction are the same, then at the equivalence point, no Sn^{2+} or I_2 remain; there is only Sn^{4+} and I^-. However, if, for example, there is an excess of Sn^{2+}, then there will be Sn^{4+}, I^- and any leftover Sn^{2+} at the equivalence point of the reaction.

FIGURE 1 Wine contains ethanol, which can be analysed by redox titrations.

10.4
PRACTICAL: CONTROLLED EXPERIMENT
What is the concentration of vitamin C in orange juice? Go to page **520**.

The end point is a change in indicator colour

Remember that the point at which an indicator changes colour is called the end point. It is important that the indicator changes colour as close as possible to the equivalence point.

self-indicating
changes colour at the equivalence point, without the need of an additional indicator

Some redox reactions do not require indicators as the species involved in the reaction will change colour at the equivalence point; they are **self-indicating**. The colour-changing species is usually used as the standard solution.

Common primary standards used in redox titrations include sodium oxalate ($Na_2C_2O_4$), potassium dichromate ($KMnO_4$), potassium bromate ($KBrO_3$), potassium dichromate ($K_2Cr_2O_7$), potassium iodate (KIO_3) and potassium hydrogen iodate ($KH(IO_3)_2$).

Colour-changing reduction half-reactions

To analyse an organic compound that can undergo oxidation, self-indicating, strong oxidising agents can be used. They cause the analyte to become oxidised, while they themselves are reduced.

Two common colour-changing oxidising agents are permanganate (which is reduced to manganese(II) ions) and dichromate (which is reduced to chromium(III) ions).

The reduction half-reaction for acidified permanganate is:

$$\underset{\text{purple}}{MnO_4^-(aq)} + 8H^+(aq) + 5e^- \rightarrow \underset{\text{clear}}{Mn^{2+}(aq)} + 4H_2O(l)$$

The purple permanganate is added to a reducing agent, which immediately converts it into the clear Mn^{2+} ion. As soon as there is no more reducing agent remaining, any additional permanganate is in excess and the solution turns dark purple (Figure 2a).

The reduction half-reaction for acidified dichromate is:

$$\underset{\text{yellow}}{Cr_2O_7^{2-}(aq)} + 14H^+(aq) + 6e^- \rightarrow \underset{\text{green}}{2Cr^{3+}(aq)} + 7H_2O(l)$$

The yellow dichromate is added to a reducing agent, which immediately converts it into the green Cr^{3+} ion. As soon as there is no more reducing agent remaining, any additional dichromate is in excess and the solution turns yellow (Figure 2b). This is a harder end point to observe due to the similarity in colour between green and yellow. Therefore, permanganate is more commonly used in redox titrations.

In the tin and iodine example we looked at before, a starch indicator can be used. Starch will turn dark blue in the presence of I_2 but is clear in the presence of I^-. As the I_2 is added to the reducing agent Sn^{2+}, it is immediately reduced to I^-, and solution remains colourless. As soon as there is no more Sn^{2+} remaining, the I_2 will be in excess, and the starch indicator will turn dark blue (Figure 2c).

$$\underset{\text{blue with starch indicator}}{I_2(aq)} + 2e^- \rightarrow \underset{\text{clear}}{2I^-(aq)}$$

Colour-changing oxidation half-reactions

To analyse an organic compound that can undergo reduction, self-indicating reducing agents can be used. Unfortunately, most strong reducing agents used in redox titrations do not typically produce colour changes at the equivalence point.

FIGURE 2 Redox titrations against a standard solution of **a** acidified potassium permanganate ($H^+/KMnO_4$), **b** acidified potassium dichromate ($H^+/K_2Cr_2O_7$) and **c** potassium iodide (KI) with starch indicator

Examples include:

- thiosulfate ($S_2O_3^{2-}$), which is oxidised to form sulfate ions (SO_4^{2-})
- oxalic acid ($C_2H_2O_4$), which is oxidised to form carbon dioxide (CO_2)
- vitamin C or ascorbic acid ($C_6H_8O_6$), which is oxidised to form dehydroascorbic acid ($C_6H_6O_6$)
- ethanol (CH_3CH_2OH), which is oxidised to form ethanoic acid (CH_3COOH).

However, the oxidation of a transition metal can produce colour changes depending on the oxidation state of the metal and the anions that it is bonded to. An example is:

$$\underset{\text{green}}{Fe^{2+}(aq)} \rightarrow \underset{\text{yellow-orange}}{Fe^{3+}(aq)} + e^-$$

FIGURE 3 Another example of a transition metal that produces different coloured solutions depending on their oxidation state is vanadium. V^{5+} is yellow, V^{4+} is blue, V^{3+} is green and V^{2+} is purple.

Calculations with titrations

Consider the titration of orange juice to determine its vitamin C (ascorbic acid) content. The orange juice is diluted in water and reacted with a standard solution of acidified permanganate, according to the overall redox reaction:

$$5C_6H_8O_6(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5C_6H_6O_6(aq) + 2Mn^{2+}(aq) + 8H_2O(l)$$

The standard solution of permanganate has a highly accurate concentration, and a known volume is reacted with the orange juice. Therefore, the amount, in mol, of permanganate can be calculated using the formula:

$$n = c \times V$$

The stoichiometric ratio in the balanced chemical equation can then be used to calculate the amount of vitamin C. If the volume of orange juice is also accurately known, the concentration of the vitamin C in the juice can be determined. Worked examples 10.4A and 10.4B demonstrate the process for the analysis of propan-1-ol and wine.

FIGURE 4 Titrations can be used to determine the vitamin C content of orange juice.

Study tip

When completing stoichiometry problems, always calculate the amount, in mol, of the chemical with known values first.

Study tip

Volume must be converted to litres in stoichiometry problems.

10.4A WORKED EXAMPLE

CALCULATING THE CONCENTRATION OF ETHANOL IN WINE USING A REDOX TITRATION

A sample of wine is analysed to calculate the concentration of ethanol (C_2H_5OH). A 25.00 mL aliquot of the wine is placed in the conical flask to react with a standard acidified solution of 1.02 M potassium permanganate to form ethanoic acid (CH_3COOH). An average titre of 25.5 mL of permanganate is measured. Determine the concentration, in M, of the ethanol.

Solution

Think	Do
Step 1: Write the half-equations for the reaction.	$C_2H_6O(aq) \rightarrow C_2H_4O_2(aq) + 4H^+(aq) + 4e^-$ $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$
Step 2: Write a balanced overall reaction equation.	$5C_2H_6O(aq) + 4MnO_4^-(aq) + 12H^+(aq) \rightarrow 5C_2H_4O_2(aq) + 4Mn^{2+}(aq) + 11H_2O(l)$
Step 3: Calculate the amount, in mol, of the 'known' chemical permanganate.	$n(MnO_4^-) = c \times V$ $= 1.02 \times (25.5 \div 1000)$ $= 0.02601\ \text{mol}$
Step 4: Use the stoichiometric ratio of the coefficients in the balanced equation (MnO_4^- : C_2H_6O is 2 : 5) to calculate the amount, in mol, of the 'unknown' chemical vitamin C.	In the 25.00 mL aliquot: $n(C_2H_6O) = \frac{5}{4} \times n(MnO_4^-)$ $= \frac{5}{4} \times 0.0261$ $= 0.03251\ \text{mol}$
Step 5: Calculate the concentration of $C_6H_8O_6$ in the 25.00 mL solution in M.	$c(C_2H_6O) = \frac{n}{V}$ $= \frac{0.03251}{25.00 \div 1000}$ $= 1.30\ \text{M}$ (3 sig fig)

10.4B Worked example
Find me in your obook pro

10.4B Worked example
Video demonstration

10.4 Skill drill
Find me in your obook pro

More steps are involved when more complex concentration units are required or solutions have been diluted as shown in Worked example 10.4B. If you're feeling confident, give Challenge 10.4 a go or practise your key science skills using Skill drill 10.4.

10.4 CHALLENGE

Calculating the concentration of vitamin C in a supplement

Some students were set the task of determining the concentration of vitamin C in nutritional supplements. A 1.0 g tablet was crushed and dissolved in 50 mL of water. This was transferred to a 250 mL volumetric flask made to the mark with water. A pipette was used to transfer 20.00 mL of diluted tablet to a 100 mL conical flask. The vitamin C was titrated with a 0.0010 M I_2 standard solution until three concordant titres are obtained. The results were 20.62 mL, 20.45 mL, 20.44 mL, 20.48 mL and 20.53 mL.

Calculate the concentration in %(*m*/*m*) (mass in 100 g) of vitamin C in the nutrient tablet, to the correct number of significant figures.

10.4 CHECK YOUR LEARNING

Describe and explain

1 Explain how redox titrations both can and cannot be self-indicating. Provide examples of both.

2 Identify one example of an organic compound that can be analysed by redox titration using permanganate solution.

3 Identify one example of an organic compound that can be analysed by redox titration using oxalic acid solution.

Apply, analyse and compare

4 Write the half-equations for the oxidation of:

- **a** thiosulfate ($S_2O_3^{2-}$) to sulfate ions (SO_4^{2-})
- **b** oxalic acid ($C_2H_2O_4$) to carbon dioxide (CO_2)
- **c** ethanol (CH_3CH_2OH) to ethanoic acid (CH_3COOH).

5 In a redox titration, 15.20 mL of acidified 0.0500 mol L^{-1} $K_2Cr_2O_7$(aq) was used to oxidise a 10.00 mL aliquot of a Sn^{2+} solution to Sn^{4+}.

- **a** Write the oxidation half-equation.
- **b** Write the reduction half-equation.
- **c** Write an overall redox equation for the reaction.
- **d** Calculate the concentration of the Sn^{2+}(aq) ions in the solution.

6 An oxalic acid solution ($H_2C_2O_4$) is made by dissolving 1.500 g in a 100.00 mL volumetric flask and used to verify that the concentration of a potassium permanganate solution is 0.0200 M. A 10.00 mL aliquot of the oxalic acid solution is titrated against an acidified solution of potassium permanganate ($KMnO_4$), resulting in an 18.26 mL average titre. The products are Mn^{2+} and CO_2 gas.

- **a** Write the oxidation half-equation.
- **b** Write the reduction half-equation.
- **c** Write an overall redox equation for the reaction.
- **d** Calculate the concentration of the $KMnO_4$ solution.
- **e** Compare the concentration of the permanganate with its hypothesised concentration and account for any differences between the values.

FIGURE 5 The ability for solutions to change colour is useful for titrations.

CHAPTER 10 Review

Chapter summary

10.1

- The bromine test is an addition reaction that uses (dark red) Br_2 that breaks double bonds in unsaturated molecules and becomes clear.
- Alcohols can be oxidised using either permanganate or dichromate. Permanganate will turn from purple to clear and dichromate will turn from yellow to green. Primary alcohols are oxidised to carboxylic acids, secondary alcohols are oxidised to ketones and tertiary alcohols cannot undergo oxidation.
- Carboxylic acids can undergo several reactions. As they are weak acids, they can react with bases to neutralise solutions. This can be shown with indicators. They also react with metals to form hydrogen gas and can undergo esterification reactions with alcohols in the presence of a sulfuric acid catalyst.

10.2

- The identity and purity of chemicals can be determined using melting and boiling points. Pure solids will have melting points within a narrow range. Impure ones have larger temperature ranges that they melt between.
- Liquids can be boiled to vaporise them into gases. Weaker intermolecular forces result in lower boiling points. Simple and fractional distillation can be used to vaporise substances and separate them into temperature fractions.

10.3

- Titrations are a type of volumetric analysis used to determine the concentration of a solute in a solvent.
- The number of double bonds in a hydrocarbon can be determined using an addition reaction with iodine (I_2). This is a quantitative technique as the number of double bonds or degree of unsaturation can be calculated.

10.4

- Redox titrations are titrations where the chemicals used will accept and donate electrons in a redox reaction. Some redox reactions change colour at their equivalence point, resulting in their ability to self-indicate their end point, rather than having to add an indicator. Others, such as iodine, do not, and an indicator such as starch must be added.

Key formulas

Concentration ($mol\,L^{-1}$)	$c = \frac{n}{V}$
Concentration ($g\,L^{-1}$ or $mg\,L^{-1}$)	$c = \frac{m}{V}$
Amount in mol from mass	$n = \frac{m}{M}$
Degree of unsaturation	$\text{Degree of unsaturation} = \frac{n(I_2)}{n(\text{hydrocarbon})}$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept: 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to go revisit them.

I can...	Confidently	Mostly	Not really	Revision link
• describe how qualitive tests can be used to determine the presence of carbon–carbon double bonds, hydroxyl and carboxyl functional groups	☐	☐	☐	Go back to Topic 10.1
• describe factors that affect the melting and boiling points and describe how melting points can indicate chemical purity	☐	☐	☐	Go back to Topic 10.2
• distinguish between fractional and simple distillation and their uses	☐	☐	☐	Go back to Topic 10.2
• calculate the degree of unsaturation from iodine titrations	☐	☐	☐	Go back to Topic 10.3
• calculate unknown concentrations using redox titrations	☐	☐	☐	Go back to Topic 10.3

Revision questions

Multiple choice

1 A primary standard should not:

A be highly impure.
B have a known chemical formula.
C have a high molar mass.
D be stable or unreactive.

2 Identify the strongest intermolecular force experienced by ammonia, NH_3.

A Dispersion forces
B Hydrogen bonding
C Covalent bonding
D Dipole–dipole attractions

3 Which set of molecules has the correct order of melting points from lowest to highest?

A C_8H_{18}, C_4H_{10}, C_2H_6
B CH_3COOH, CH_3CH_2OH, CH_3CH_3
C CH_3OH, $CH_3CH_2CH_2OH$, CH_3CH_2OH
D CH_3COOH, CH_3CH_2COOH, $CH_3(CH_2)_2COOH$

4 Which of the following is the most accurate indicator of the end point of a titration?

A Temperature increase
B Temperature decrease
C Colour change
D Bubbling of solution

5 What is the correct definition of a titre?

A A graduated glass cylinder that dispenses highly accurate volumes of a standard solution
B The solution in the burette that is added to the solution in the conical flask
C The volume of solution dispensed from the burette
D A glass tube that dispenses highly accurate volumes of a solution

6 To determine the titre volume, you have to:

A subtract the initial volume from the final volume to determine the total volume of solution dispensed from the burette.
B subtract the final volume from the initial volume to determine the total volume of solution dispensed from the burette.
C subtract the initial volume from the final volume to determine the total volume of solution dispensed from the pipette.
D subtract the final volume from the initial volume to determine the total volume of solution dispensed from the pipette.

7 The copper(II) ions in a solution can be converted to copper metal by trickling the solution over scrap iron to produce an iron(II) solution. If the process produces 15.50 L of solution containing 0.003 50 mol L^{-1} of Fe^{2+}(aq) ions, what mass of copper is produced?

- **A** 3.44 g
- **B** 1.72 g
- **C** 3.03 g
- **D** 1.51 g

The following information refers to Questions 8–10.

In a redox titration, the concentration of ascorbic acid in a lemon was determined by squeezing all the juice from the lemon and filtering it to remove the pulp. A 10.00 mL sample of the juice was added to a 250.00 mL volumetric flask, which was made to the mark with distilled water. 25.00 mL aliquots of this solution were titrated against a solution of 1.64×10^{-4} M iodine solution with a starch indicator. An average titre of 14.60 mL was measured.

$$C_6H_8O_6(aq) + I_2(aq) \rightarrow C_6H_6O_6(aq) + 2I^-(aq) + 2H^+(aq)$$

8 Calculate the concentration of ascorbic acid in lemon juice in g L^{-1}.

- **A** 0.0024
- **B** 0.43
- **C** 0.024
- **D** 0.043

9 Which of the following errors would result in an underestimation of the concentration of the standard solution?

- **A** The conical flask was not dried and had approximately 1 mL of water in it when the aliquot was added.
- **B** The initial titre volume was read at eye level to the meniscus, but the final volume was read from above the meniscus.
- **C** The concentration of the iodine solution was miscalculated, and was actually 1.62×10^{-4} M.
- **D** The volumetric flask used was 500.00 mL rather than 250.00 mL.

10 Which of the following errors would result in an overestimation of the concentration of the solution?

- **A** The pipette was rinsed with iodine solution.
- **B** The pipette was rinsed with the diluted lemon juice solution.
- **C** The conical flask was rinsed with water.
- **D** The conical flask was rinsed with diluted lemon juice solution.

Short answer

Describe and explain

11 Explain two chemical tests that can identify the presence of a carbon–carbon double bond. In your answer you must provide:

- **a** chemical equations
- **b** the observations – changes that would be seen if the chemical had carbon–carbon double bonds
- **c** the limitation of the technique in determining the identity of a molecule.

12 In a chemical test, explain how you would determine whether an alcohol was primary, secondary or tertiary. In your answer, you must provide:

- **a** balanced chemical equations
- **b** the observations – changes that would be seen if the chemical was a primary, secondary or tertiary alcohol
- **c** the limitations of the technique in determining the identity of a primary, secondary or tertiary alcohol.

13 Explain three chemical tests that could be used to identify the presence of a carboxylic acid. In your answer you must provide:

- **a** balanced chemical equations
- **b** the observations – changes that would be seen if the chemical was a carboxylic acid
- **c** the limitation of the technique in determining the identity of a carboxylic acid.

14 Explain the difference in melting points of sodium chloride (NaCl) 801°C, water (H_2O) 0°C and methane (CH_4) –182°C.

15 Explain whether substances with higher boiling points have stronger or weaker intermolecular forces.

16 Explain how a substance's purity can affect its melting point.

17 Differentiate between the terms *aliquot* and *titre* and identify the solutions they refer to in a titration.

18 Explain why a solution of permanganate is more commonly used than a solution of dichromate in a redox titration.

FIGURE 1 **a** Permanganate and **b** dichromate have distinct colours.

19 Explain the effect of only washing a burette with water on the final concentration of an unknown solution in a redox titration.

20 Explain the effect of leaving water in the conical flask on the final concentration of an unknown solution in a redox titration.

Apply, analyse and compare

21 Compare the intermolecular forces of ethanol and water and explain why ethanol has a lower boiling point than water.

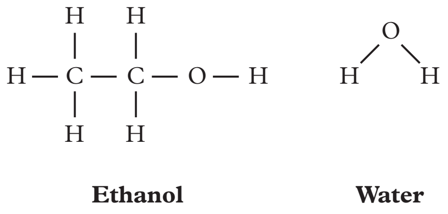
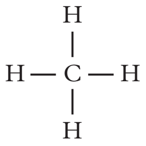


FIGURE 2 Structures of ethanol and water

22 Compare the structures and intermolecular forces in methane (CH_4) and octane (C_8H_{18}). Explain why methane has a boiling temperature of –161.5°C and is a gas at room temperature, whereas octane has a boiling temperature of 125.7°C and is a liquid at room temperature.

23 Determine which molecule in the following pairs of molecules would have a higher boiling point. Explain your reasoning.

a



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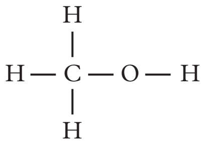
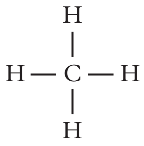


b



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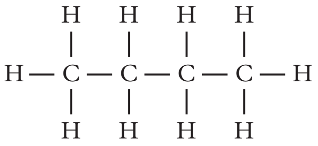
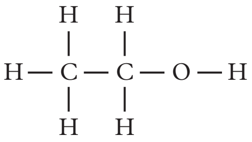


c



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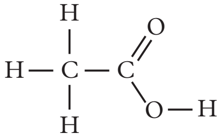

24 A 0.0450 M $KMnO_4$ solution is used to oxidise all of the Br^- ions in 25.00 mL of an acidic 0.200 M NaBr(aq) solution.

a Write the oxidation half-equation.

b Write the reduction half-equation.

c Write an overall redox equation for the reaction.

d Calculate the volume of $KMnO_4$ solution required in the titration.

25 A bottle of betadine claims to contain 10%(*m*/*v*) of iodine. To determine the accuracy of this claim, 10.00 mL aliquots of betadine were titrated against a 0.15 M solution of thiosulfate. In the reaction, thiosulfate ($S_2O_3^{2-}$) is oxidised to tetrathionate ($S_4O_6^{2-}$) and iodine (I_2) is reduced to iodide (I^-). The average titre of thiosulfate was 25.74 mL.

a Write a half-equation for thiosulfate.

b Write a half-equation for iodine.

c Write an overall redox equation for the reaction.

d Calculate the amount, in mol, of thiosulfate that reacted.

e Calculate the amount, in mol, of iodine in the 10 mL aliquot.

f Calculate the concentration of iodine in betadine in %*(m/v)*.

FIGURE 3 Betadine is an antiseptic used to disinfect skin for surgery.

26 The mass of iron in a 10.0 g sample of iron ore is determined by redox titration with an acidified standard solution of 0.055 M potassium permanganate ($KMnO_4$). The 10.0 g sample is completely dissolved in a 250 mL volumetric flask to form Fe^{2+}. The average titre obtained from a 10.00 mL aliquot is 17.26 mL. The products of the reaction are Fe^{3+} and Mn^{2+}.

a Write the oxidation half-equation.

b Write the reduction half-equation.

c Write an overall redox equation for the reaction.

d Calculate the mass of iron in the 10 mL sample of ore.

e Calculate the %*(m/m)* of iron in the ore.

27 Sulfur dioxide is a preservative used in wine to prevent the wine from oxidising. Most manufacturers have SO_2 concentrations between 10 and 20 ppm ($mg\,L^{-1}$) although some may go up as high as 400 ppm, especially in sweeter dessert wines.

To measure its concentration, NaOH is added to the wine to convert all sulfites to SO_2, and then H_2SO_4 is added to convert the aqueous SO_2 into a gas. A redox titration is then performed where $SO_2(g)$ is converted to $SO_4^{2-}(aq)$ using a 4.85×10^{-3} M iodine (I_2) solution and a starch indicator. A 25.00 mL aliquot of wine, from a 750 mL bottle, is placed in a conical flask and an average titre volume of 17.40 mL is calculated.

a Write the oxidation half-equation.

b Write the reduction half-equation.

c Write an overall redox equation for the reaction.

d Calculate the amount, in mol, of I_2.

e Calculate the amount, in mol, of SO_2 in the 25.00 mL aliquot.

f Calculate the concentration of SO_2 in ppm.

FIGURE 4 Red wine

28 30 vol hair dye has a peroxide (H_2O_2) concentration of 8.8%(*m/v*). In solution, hydrogen peroxide will naturally decompose into water and oxygen gas. To verify the concentration of peroxide in hair dye, 10.00 mL aliquots are titrated against 0.800 mol L^{-1} acidified $K_2Cr_2O_7$ solution. The products of the reaction are $H_2O(l)$, $O_2(g)$ and $Cr^{3+}(aq)$. For an average titre volume of 11.87 mL:

a write the oxidation half-equation

b write the reduction half-equation

c write an overall redox equation for the reaction

d calculate the concentration of peroxide in %(*m*/*v*).

Design and discuss

29 Design a series of experiments and place them into a flowchart to determine whether an unlabelled chemical was a primary alcohol or a carboxylic acid. For each step, state what you are testing for and how you would know if the test were positive.

30 Design a series of experiments and place them into a flowchart to determine whether an unlabelled chemical was a saturated or unsaturated hydrocarbon. For each step, state what you are testing for and how you would know if the test were positive.

31 The boiling point of 1-propanol is 97°C and that of ethyl methyl ether is 7.4°C. The two molecules have the same number of carbon, oxygen and hydrogen atoms (Figure 5). Analyse the molecules and justify the difference in their boiling points.

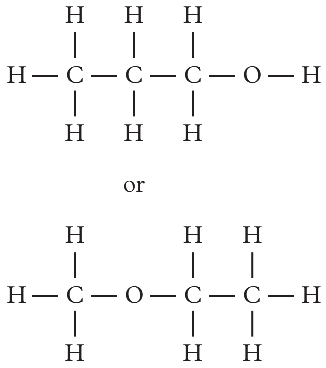

FIGURE 5 Structures of 1-propanol and ethyl methyl ether

32 Discuss the effect of the following errors on the final concentration calculated in a redox titration.

a The initial titre volume is read from below the meniscus; the final is read with the meniscus at eye level.

b A bubble comes out of the burette during the titration.

c A standard solution has a concentration higher than what it actually is.

d A standard solution has a concentration lower than what it actually is.

e The pipette is rinsed with water but not with the aliquot solution.

f The burette is rinsed with water but not with the titrant solution.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER 11

Instrumental analysis of organic compounds

KEY KNOWLEDGE

- applications of mass spectrometry (excluding features of instrumentation and operation) and interpretation of qualitative and quantitative data, including identification of molecular ion peak, determination of molar mass and identification of simple fragments
- identification of bond types by qualitative infrared spectroscopy (IR) data analysis using characteristic absorption bands
- structural determination of organic compounds by low resolution carbon-13 nuclear magnetic resonance (^{13}C-NMR) spectral analysis, using chemical shift values to deduce the number and nature of different carbon environments
- structural determination of organic compounds by low and high resolution proton nuclear magnetic resonance (^{1}H-NMR) spectral analysis using chemical shift values, integration curves (where the height is proportional to the area underneath a peak) and peak splitting patterns (excluding coupling constants), and application of the n+1 rule (where n is the number of neighbouring protons), to deduce the number and nature of different proton environments)
- the principles of chromatography including use of high performance liquid chromatography (HPLC) and retention times and the construction of a calibration curve to determine the concentration of an organic compound in a solution (excluding features of instrumentation and operation)
- deduction of the structures of simple organic compounds using a combination of mass spectrometry (MS), infrared spectroscopy (IR), proton nuclear magnetic resonance (^{1}H-NMR) and carbon-13 nuclear magnetic resonance (^{13}C-NMR) (limited to data analysis)
- the roles and applications of laboratory and instrumental analysis, with reference to product purity and the identification of organic compounds or functional groups in isolation or within a mixture

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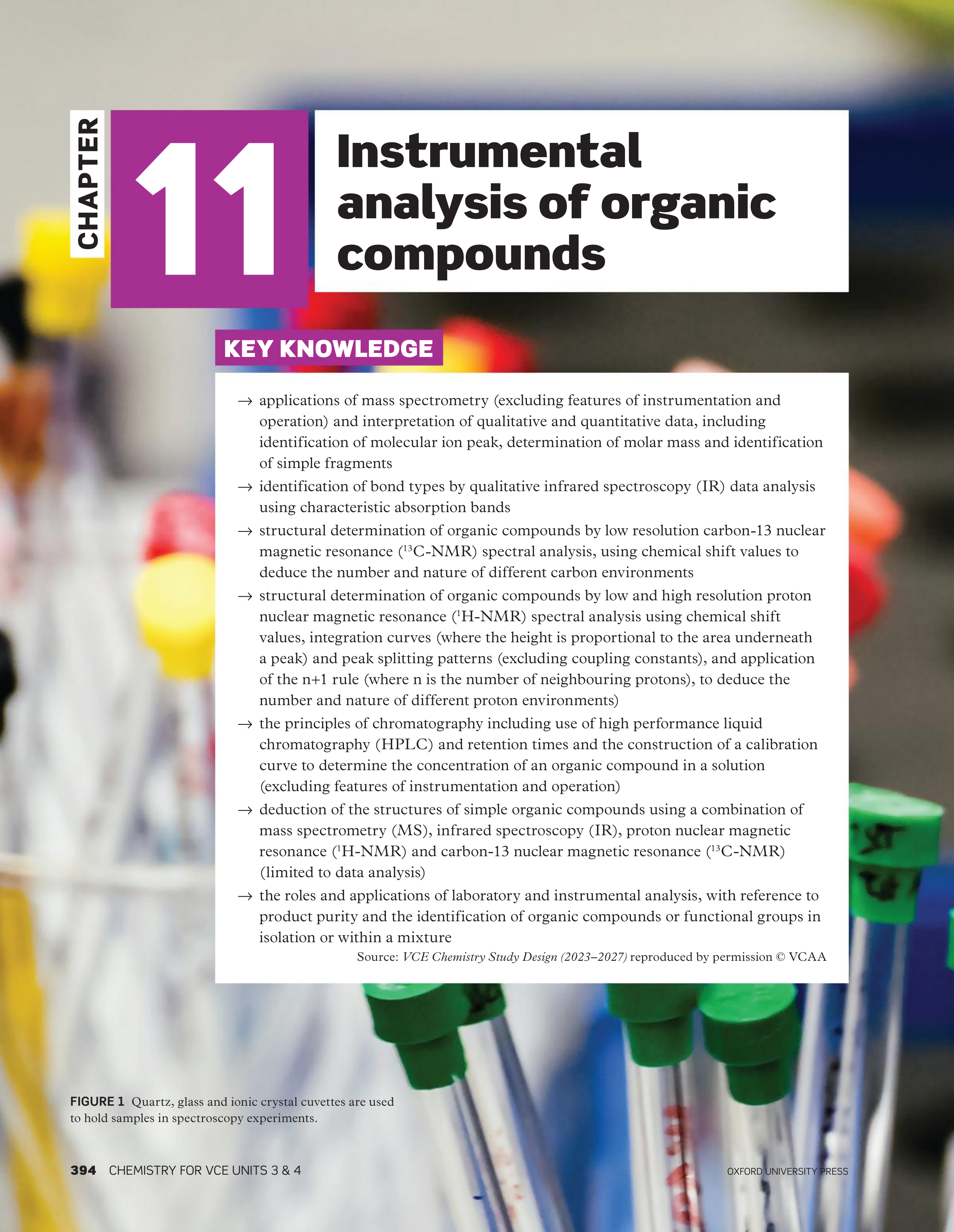

FIGURE 1 Quartz, glass and ionic crystal cuvettes are used to hold samples in spectroscopy experiments.

GROUNDWORK

In Chapter 11, you will learn about how structures of organic molecules can be determined using different analytical techniques, such as nuclear magnetic resonance spectroscopy, mass spectrometry and high performance liquid chromatography, and infrared spectroscopy.

This chapter will build on concepts from Chapters 8, 9 and 10. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

11A Describe an organic compound.

11A Groundwork resource
Organic compounds

11B Identify the types of intramolecular bonds in organic molecules.

11B Groundwork resource
Intramolecular bonding

11C Describe the types of analyses that can be done to identify organic compounds.

11C Groundwork resource
Qualitative analysis of organic compounds

11D Contrast quantitative and qualitative analysis.

11D Groundwork resource
Quantitative and qualitative analysis

PRACTICALS

11.5A **PRACTICAL: CASE STUDY**	How can analytical techniques be used to differentiate between isomers?	Page 522
11.5B **PRACTICAL: CASE STUDY**	Lab clean-up duty – what's in the bottle?	pro
11.7 **PRACTICAL: FIELDWORK**	How is instrumental analysis used in industry?	Page 523

11.1 Mass spectrometry

KEY IDEAS

In this topic, you will learn that:

- the molar mass of a compound can be determined using mass spectrometry
- a compound can be identified based on how it fragments in a mass spectrometer.

mass spectrometry an analytical technique in which the masses and relative abundances of isotopes, molecules and molecular fragments are measured by ionising them and determining their trajectories in electric and magnetic fields

In Unit 1 of VCE Chemistry, you looked at identifying the relative atomic mass of an element using **mass spectrometry** (MS). This topic will look at how different organic compounds can be identified from their unique mass spectra obtained from a mass spectrometer.

Mass spectrometry

Mass spectrometers work differently from the other spectroscopy instruments that we will look at in this chapter. In a mass spectrometer:

- the sample is vaporised with a heater and bombarded by high-energy electrons
- these electrons have enough energy to knock electrons off atoms and create positively charged ions, which are accelerated through a magnetic field that separates ions based on their mass-to-charge ratio (m/z)
- the ions that move more slowly (i.e. the heavier ions) are deflected less, while ions that move quickly (i.e. the lighter ions) are deflected more
- the different movements of the ions are then analysed by the detector, which produces a unique mass spectrum for the sample.

Study tip

You do not need to memorise how a mass spectrometer works, only how to use the spectra obtained to characterise organic compounds.

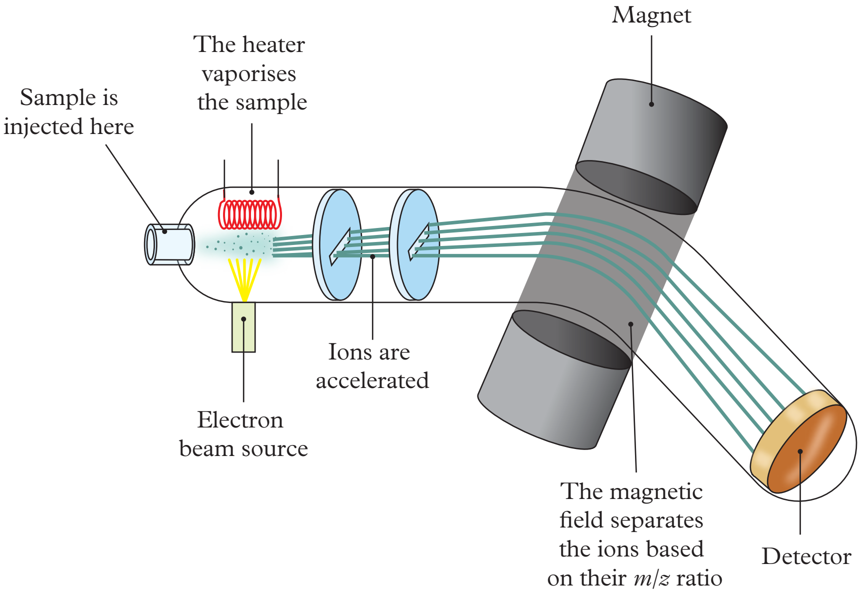

FIGURE 1 Simplified diagram of a mass spectrometer

Analysing mass spectra

Inside the mass spectrometer, an organic compound can be fragmented into many different positively charged ions. These ions are deflected at different angles depending on their **mass-to-charge ratio (*m*/*z*)**. Each ion produces a different peak at a unique *m*/*z* value, which is plotted on the *x*-axis of a **mass spectrum**. The *y*-axis shows the **relative intensity** or abundance of each of the ions that fragment from the compound.

Let's look at ethanoic acid, CH_3COOH, as an example. In the mass spectrometer, ethanoic acid is bombarded by electrons and ionised, with an electron removed, as follows:

$$CH_3COOH + e^- \rightarrow CH_3COOH^+ + 2e^-$$

CH_3COOH^+ is the **molecular ion**. It is unstable and can break down into many other **fragmented ions**, as shown in Figure 2.

If you look at the mass spectrum of ethanoic acid in Figure 3, you can see two annotated peaks. One is the **molecular ion peak**, which occurs when the whole original molecule becomes positively charged. In this spectrum, the unfragmented molecular ion is $[CH_3COOH]^+$, which has a *m*/*z* of 60. For this ion, *m*/*z* is equal to molar mass because it has a charge of +1, and the molar mass of $[CH_3COOH]^+$ is $60\,g\,mol^{-1}$.

mass-to-charge ratio (*m*/*z*)
the molar mass of the positive ion formed divided by its charge

mass spectrum
a column graph that shows the relative abundance of each ionised fragment in a sample

relative intensity/abundance
the percentage of ions produced by the mass spectrometer

molecular ion
the unfragmented, ionised (positively charged) form of the original molecule, which produces the peak on a mass spectrum

fragmented ions
the smaller pieces formed when the bonds holding the molecular ion together are broken

molecular ion peak
the peak on the mass spectrum that corresponds to the positively charged, unfragmented, whole, original molecule

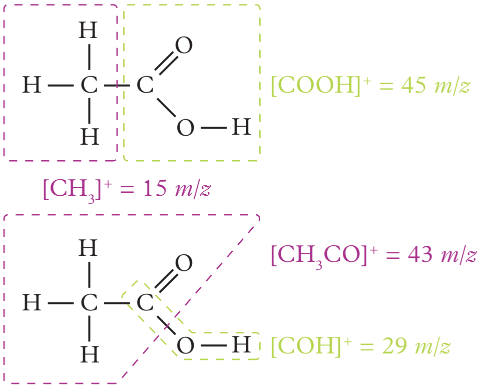

FIGURE 2 Some of the fragmented ions for ethanoic acid

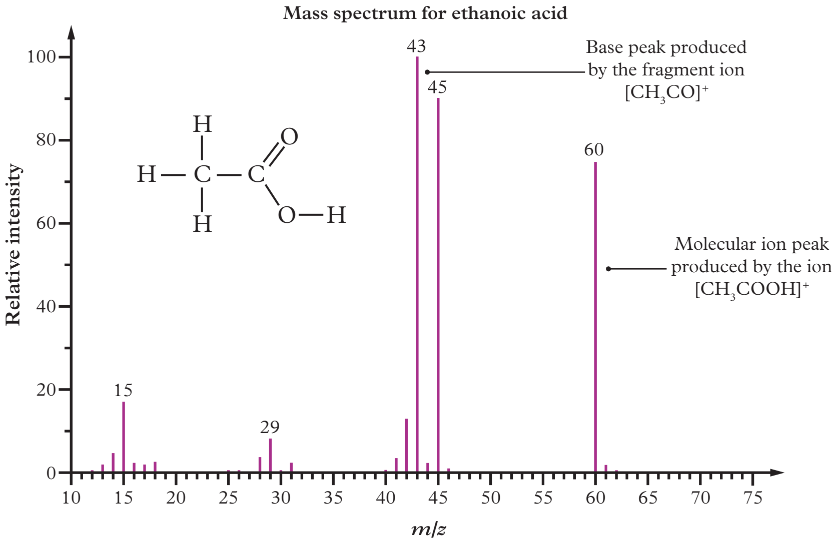

FIGURE 3 Mass spectrum of ethanoic acid

The second annotated peak is the **base peak**. This peak represents the most intense or abundant fragment. Usually, the base peak is an ion created by a single split in the original ion. It is the most stable ion fragment formed. Fragmenting can happen at any bond within the molecule, and each peak on the mass spectrum represents a different fragment that has been ionised from the original molecule.

For ethanoic acid, the base peak occurs at *m*/*z* = 43. This means that the most abundant ion produced from the fragmenting of ethanoic acid is the ion $[CH_3CO]^+$ ($M = 43\,g\,mol^{-1}$). Can you identify which fragments of ethanoic acid correspond to the remaining peaks?

base peak
the peak on the mass spectrum that has the highest relative intensity; it is the most abundant and stable ion present

11.1 Worked example
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11.1 Worked example
Video demonstration

TABLE 1 Common fragments and their *m*/*z* ratio

Fragments	*m*/*z*
$[CH_3]^+$	15
$[OH]^+$	17
$[CH_3CH_2]^+$, $[COH]^+$	29
$[CH_2OH]^+$	31
$[CH_3CH_2CH_2]^+$, $[CH_3CO]^+$	43
$[COOH]^+$, $[CH_2CH_2OH]^+$	45
$[CH_2CH_2CH_2OH]^+$	59

The *m*/*z* ratio of common fragments

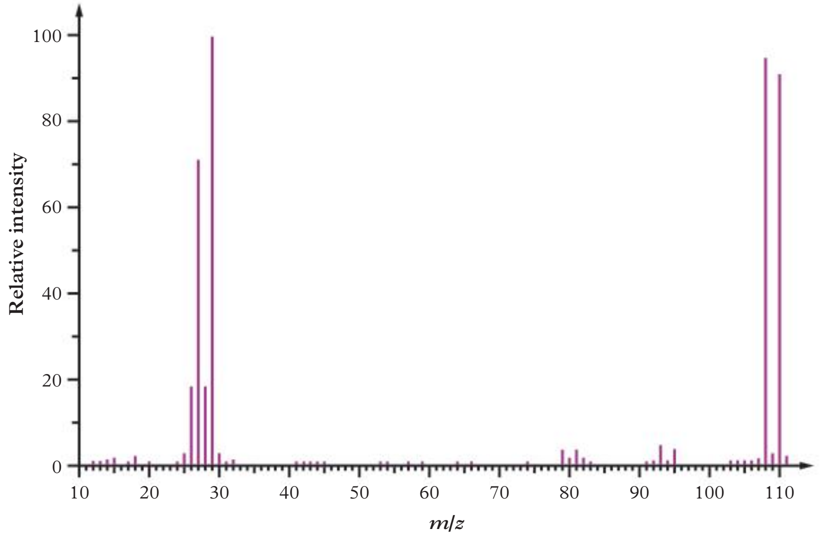
Mass spectrometry produces many different fragments. You may find it useful to remember some of the more common fragments and their *m*/*z* values (Table 1). See how you can apply your understanding of fragmentation in Worked example 11.1.

The isotopic effect

Most elements exist as different **isotopes**. This can sometimes cause additional peaks to appear in a mass spectrum.

isotopes
different forms of the same element with the same number of protons but a different number of neutrons

For example, the element bromine has two isotopes: ^{79}Br and ^{81}Br. Both isotopes have similar relative abundances of 50.7% and 49.3%, respectively. Therefore, the mass spectrum of bromoethane will have two molecular ion peaks of similar relative intensities (Figure 4).



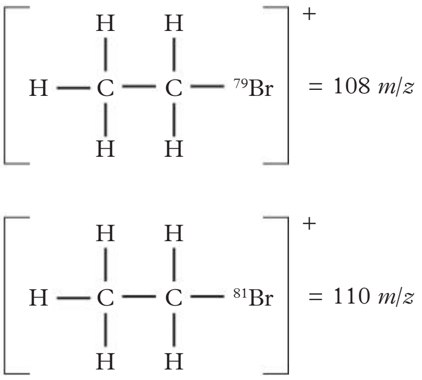


FIGURE 4 Mass spectrum of bromoethane

TABLE 2 Common isotopes and their percentage abundance

Element	Isotopes	% abundance
Hydrogen	^{1}H	99.98
	^{2}H	0.02
Carbon	^{12}C	98.9
	^{13}C	1.1
Oxygen	^{16}O	99.8
	^{18}O	0.2
Chlorine	^{35}Cl	75.8
	^{37}Cl	24.2
Bromine	^{79}Br	50.8
	^{81}Br	49.2

There are many other elements with multiple isotopes (Table 2).

^{2}H, ^{13}C and ^{18}O are not very abundant isotopes. The peaks that these isotopes produce are very small and can be mistakenly ignored. Chlorine and bromine, on the other hand, have higher proportions of each isotope. The mass spectra of compounds containing these elements will have significant peaks to represent each isotope, as you saw in Figure 4 for bromoethane.

A compound containing elements with multiple abundant isotopes will produce more than one molecular ion peak.

Study tip

m/*z* is the same as the molar mass of the compound or fragment. So, the *m*/*z* for the molecular ion peak is the molar mass of the compound.

Study tip

Remember to assign a positive charge to each fragment produced on a mass spectrum.

For example, 1,1-dibromoethane has three molecular ion peaks (Figure 5):

- The peak at $m/z = 186$ is produced by the molecular ion containing two ^{79}Br atoms.
- The peak at $m/z = 188$ is produced by the molecular ion containing one ^{79}Br atom and one ^{81}Br atom.
- The peak at $m/z = 190$ is produced by the molecular ion containing two ^{81}Br atoms.

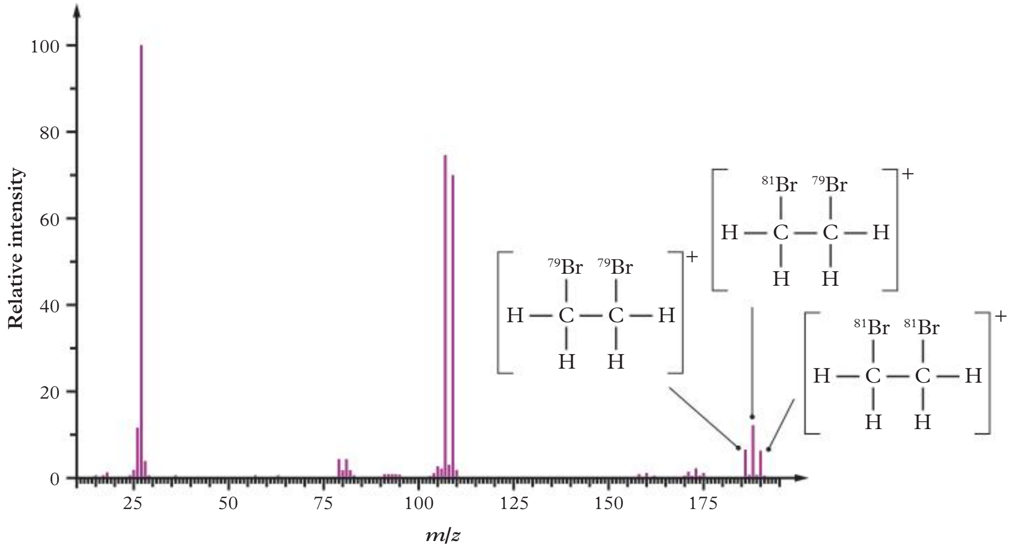


FIGURE 5 The mass spectrum of 1,1-dibromethane

11.1 CHECK YOUR LEARNING

Describe and explain

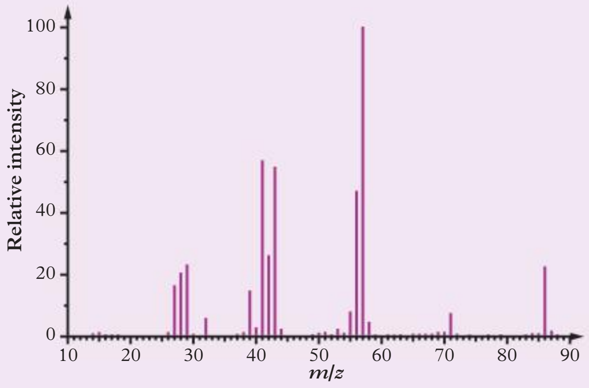
1 Describe the terms *base peak* and *molecular ion peak*.

2 **a** Explain the presence of three molecular ion peaks in the mass spectrum of 1,2-dichloroethane.

b Identify the m/z of each molecular ion peak.

Apply, analyse and compare

3 Analyse the following mass spectrum of a straight-chain alkane and answer the questions.



a Identify the molar mass of the compound.

b Identify the m/z value of the base peak.

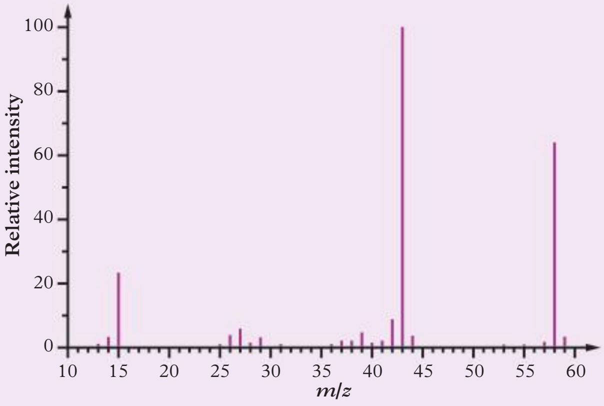
c Identify the fragment that caused the base peak.

d Identify the molecular formula of the compound.

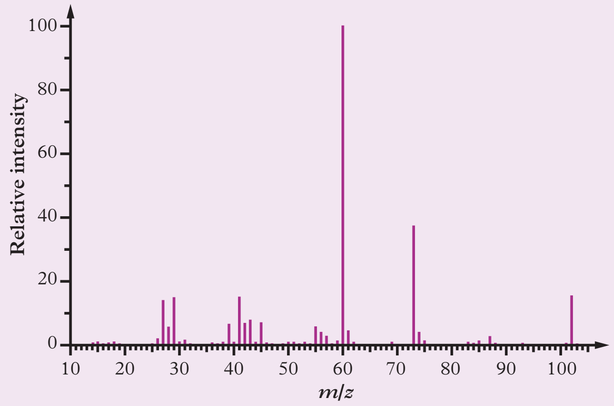
e Identify the fragment that caused the peak at $m/z = 43$.

4 Use the mass spectra shown below to identify the two compounds.

a An aldehyde

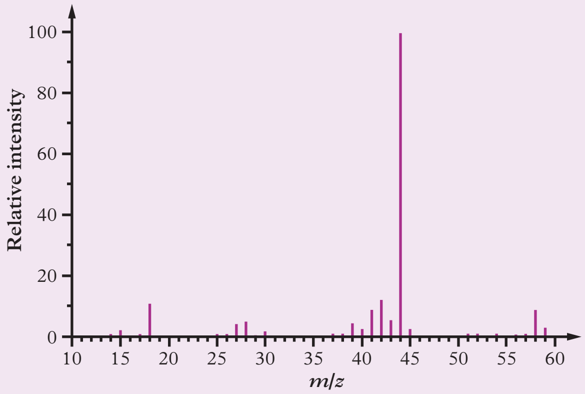


b A carboxylic acid

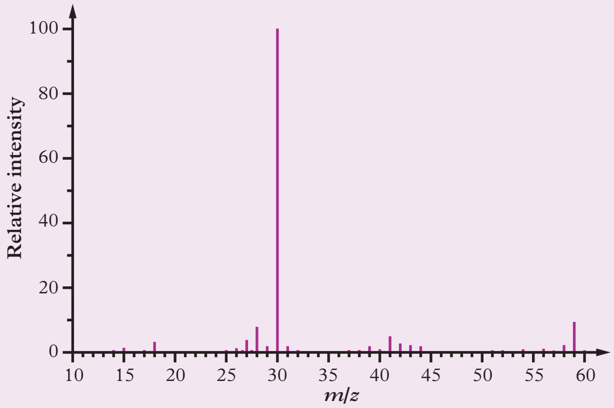


5 Two isomers of the same amine were run through a mass spectrometer. The spectra produced are shown below.

Spectrum 1



Spectrum 2

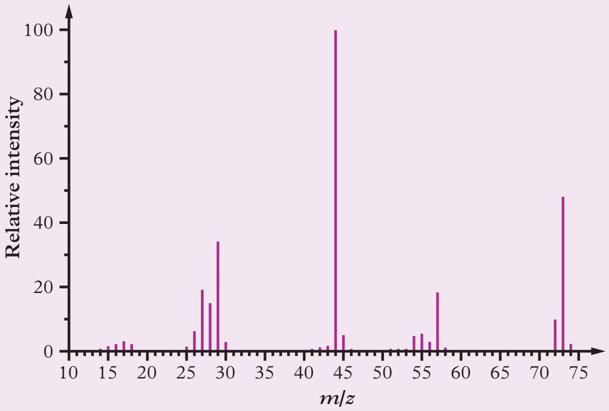


a Identify the molecular formula of the amine.

b Identify the *m*/*z* of the base peak in each spectrum.

c Identify the two isomers that caused each spectrum. Explain your reasoning.

6 A primary anide was analysed using mass spectrometry, producing the following spectrum.



a Explain the presence of the peak at *m*/*z* = 73.

b Identify the molar mass of the compound.

c Identify the ion that caused the base peak.

d Draw the structural formula of the primary amine.

Design and discuss

7 Discuss whether a mass spectrum could be used to determine if a sample was a ketone or an aldehyde with the same molecular formula. Explain your reasoning.

11.2

Infrared spectroscopy

KEY IDEAS

In this topic, you will learn that:

- the functional groups in organic compounds can be identified using infrared spectroscopy.

infrared spectroscopy an analytical technique in which bond type can be identified by measuring the amount of infrared radiation absorbed by the bonds

electromagnetic spectrum the range of wavelengths that all electromagnetic radiation exists at, including UV and visible light, infrared radiation, microwaves and radio waves

infrared radiation a wavelength region on the electromagnetic spectrum of 700–1000 nm

Organic compounds can contain many different functional groups. The bonds holding the functional groups together respond differently to infrared radiation. Because of this, qualitative **infrared spectroscopy (IR)** can be used to differentiate between the functional groups of organic compounds. When used in conjunction with other techniques such as MS (which you learnt about in the previous topic), ^{13}C-NMR and ^{1}H-NMR (which you will learn about later), it is possible to determine the full structure of an organic compound.

Infrared spectroscopy

Infrared spectroscopy uses light from the infrared region of the **electromagnetic spectrum** (Figure 1). **Infrared radiation** has wavelengths of 700–1000 nanometres (nm).

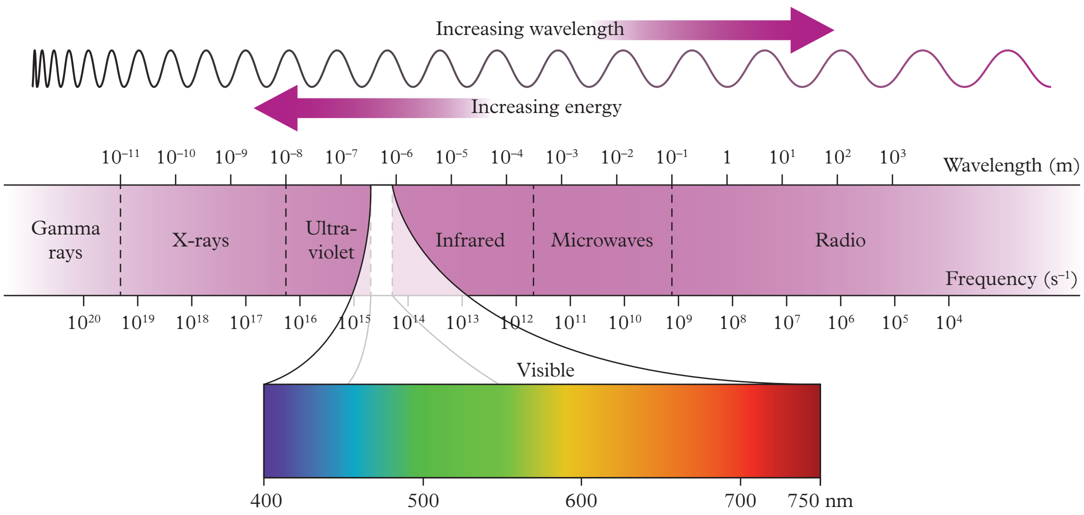


FIGURE 1 The electromagnetic spectrum

When a bond is subjected to infrared radiation, it experiences different types of vibrations and rotations. The vibrations can be classified as bending or stretching (Figure 2).

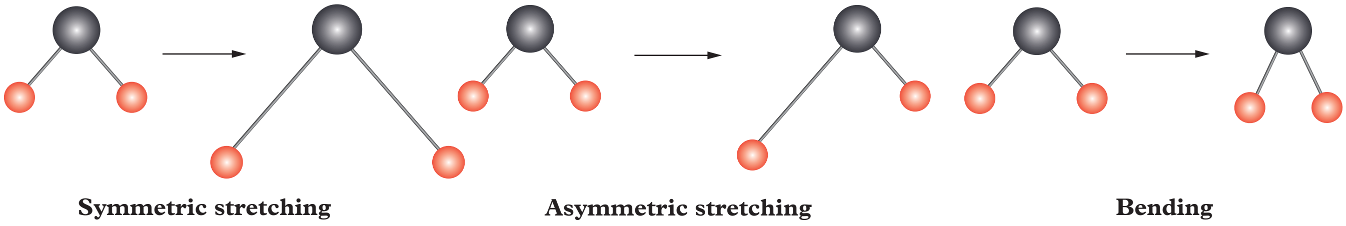


FIGURE 2 Bending and stretching in molecules

Different functional groups will consist of different bond types. These bonds will absorb specific wavelengths of infrared radiation, giving each bond a characteristic wavelength absorption, or **wavenumber**. The wavenumbers of common bonds present in functional groups are shown in Table 1. Wavenumbers translate to peaks in an infrared spectrum. The peaks on the spectrum can then be used to identify the functional groups present in an organic compound.

wavenumber
a unit of frequency (cm^{-1})

TABLE 1 Characteristic ranges for infrared absorption

Bond	Wavenumber (cm^{-1})
C–Cl (chloroalkanes)	600–800
C–O (alcohols, esters, ethers)	1050–1410
C=C (alkenes)	1620–1680
C=O (amides)	1630–1680
C=O (aldehydes)	1660–1745
C=O (acids)	1680–1740
C=O (ketones)	1680–1850
C=O (esters)	1720–1840
C–H (alkanes, alkenes, arenes)	2850–3090
O–H (acids)	2500–3500
O–H (alcohols)	3200–3600
N–H (amines and amides)	3300–3500

Analysing IR spectra

The *y*-axis on an IR spectrum represents **transmittance** (%). This is the amount of infrared radiation passing through the sample. The greater the transmittance, the more light is passing through – 100% transmittance means that 100% of the radiation is passing through the sample. If a compound contains certain functional groups, the specific bonds will absorb infrared radiation at specific wavenumber ranges and cause the transmittance to decrease, i.e. less radiation is passing through. Because of this, the peaks appear upside down in IR spectra.

transmittance
the amount of infrared radiation passing through a sample

The *x*-axis shows the wavenumber (cm^{-1}). The wavenumber scale increases from right to left.

When analysing an IR spectrum, we look to the *x*-axis to determine the functional groups that are present in the compounds. Let's look at the IR spectrum of propanoic acid as an example (Figure 3).

functional group region
the absorption pattern in the wavenumber range above 1400 cm^{-1}; provides information about the functional groups present in the compound

fingerprint region
the absorption pattern in the wavenumber range 500–1400 cm^{-1}; unique to each compound

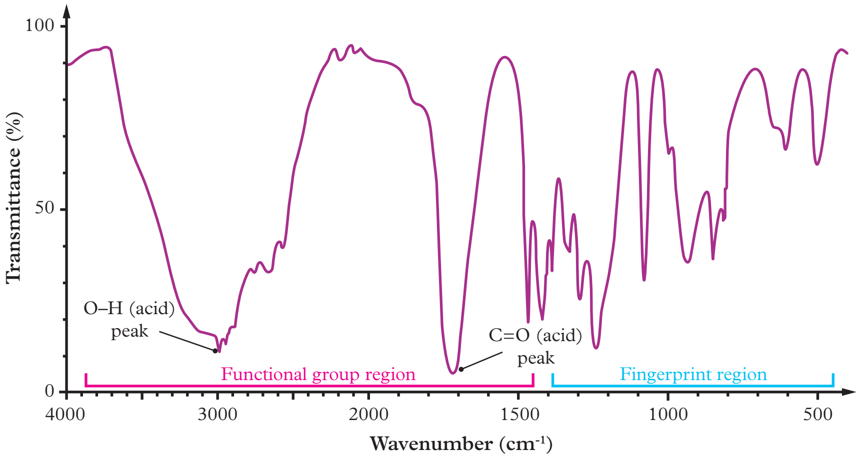


FIGURE 3 The IR spectrum of propanoic acid

Study tip

Because there are so many peaks in the fingerprint region, you are not expected to identify peaks here. If there are significant peaks or bonds in this region, they will be identified for you in exam questions.

Two peaks have been labelled: the broad peak at ~3000 cm^{-1} indicates an O–H (acid) bond and the narrower peak at ~1700 cm^{-1} is the C=O bond. They both come from the carboxyl functional group in propanoic acid. These peaks are in the **functional group region**, above 1400 cm^{-1} in IR spectra. The wavenumber ranges in which different bond types absorb infrared radiation are shown in Figure 4. There is also a **'fingerprint' region** at 500–1400 cm^{-1}. Like human fingerprints, this region is unique to each compound.

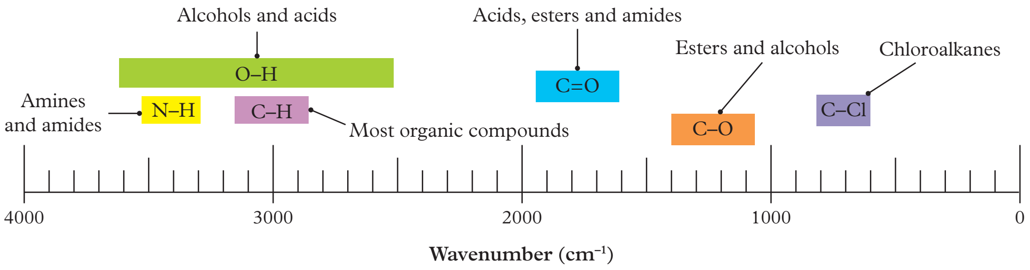


FIGURE 4 Different functional groups and where they appear on IR spectra

Many types of peaks, or **absorption bands**, can appear on an IR spectrum. They can be described as strong, medium or weak in relation to the depth of the band. Absorption bands are also described as broad or narrow in relation to the width of the band. Figure 5 shows the different types of peaks. In Figure 3, you can describe the peak at around 3000 cm^{-1} as strong and broad, and the peak at around 1700 cm^{-1} as strong and narrow.

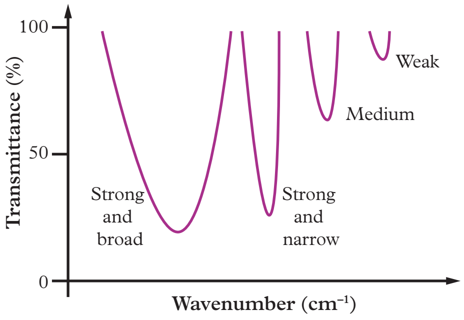


FIGURE 5 Different types of peaks that appear on an IR spectrum

absorption bands
peaks in an IR spectrum

Common spectra

To identify organic compounds from IR spectra, it is useful to become familiar with common spectra. Organic compounds in the same homologous series will have a characteristic generic shape or absorption pattern.

It is common for an alcohol to have a strong, broad peak that is very smooth, indicating the O–H bond in the 3200–3600 cm^{-1} region (Figure 6).

Carboxylic acids also have a strong, broad peak for their O–H bonds in the same wavenumber region (2500–3500 cm^{-1}) as alcohols. However, it is much more jagged and pointy than the smooth O–H peak in an alcohol (Figure 7). Carboxylic acids also have a strong, narrow peak at ~1700 cm^{-1} for their C=O, which alcohols do not have.

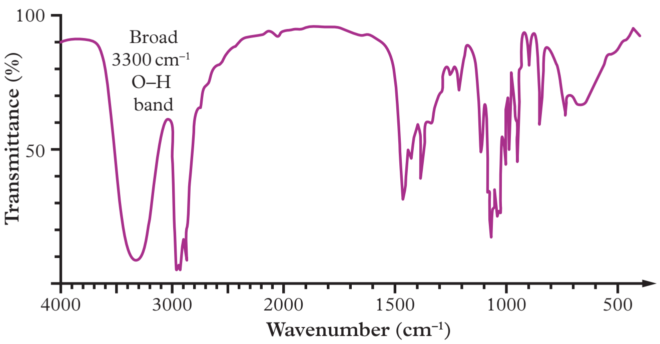


FIGURE 6 An IR spectrum of an alcohol

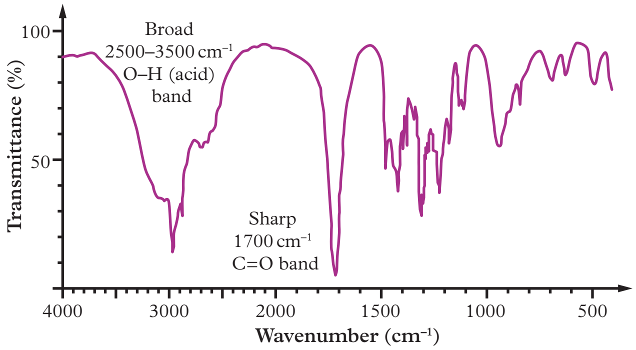


FIGURE 7 An IR spectrum of a carboxylic acid

The N–H bond that is present in amines absorbs light at a wavenumber greater than 3000 cm^{-1} on the IR spectrum. However, the N–H bond is a medium, narrow peak (Figure 8). The peak is quite pointy – almost like a fang.

An amide will have a stronger peak in the 3300–3500 cm^{-1} region, but with a similar sharp, pointy style to the N–H peak in an amine (Figure 9). It will also have a strong, narrow peak representing the C=O bond at 1630–1680 cm^{-1}.

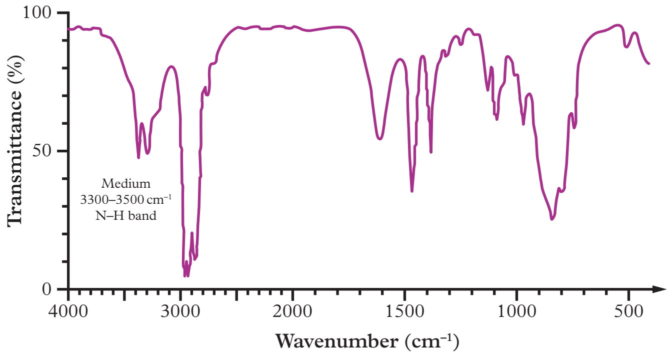


FIGURE 8 An IR spectrum of an amine

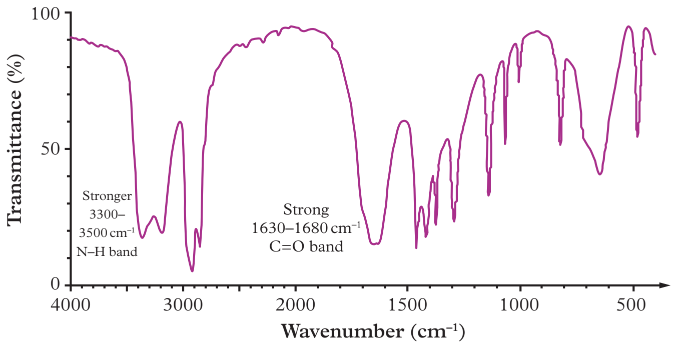


FIGURE 9 An IR spectrum of an amide

Study tip

Being able to quickly identify functional groups is important, so make sure you know what the common spectra look like.

Study tip

On the exam, you must include the units cm^{-1} when talking about wavenumbers.

Unlike the previous homologous series, you cannot look at the region above 3000 cm^{-1} alone to help identify it as an ester. This is because esters contain a C–O bond, which appears in the fingerprint region, as well as a C=O bond, which falls between 1720–1840 cm^{-1} as a strong, narrow peak (Figure 10).

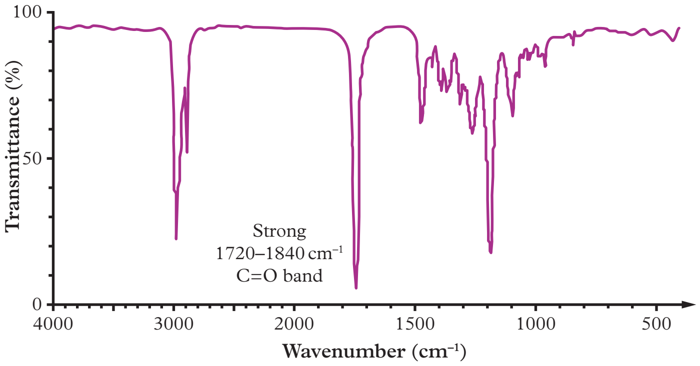


FIGURE 10 An IR spectrum of an ester

You may have also noticed that each of the spectra in Figures 6–10 has a sharp peak at 2850–3090 cm^{-1}. This corresponds to the C–H bond, which is present in most organic compounds.

Explore the application of IR in breathalysers using Real-world chemistry 11.2.

11.2 Real-world chemistry
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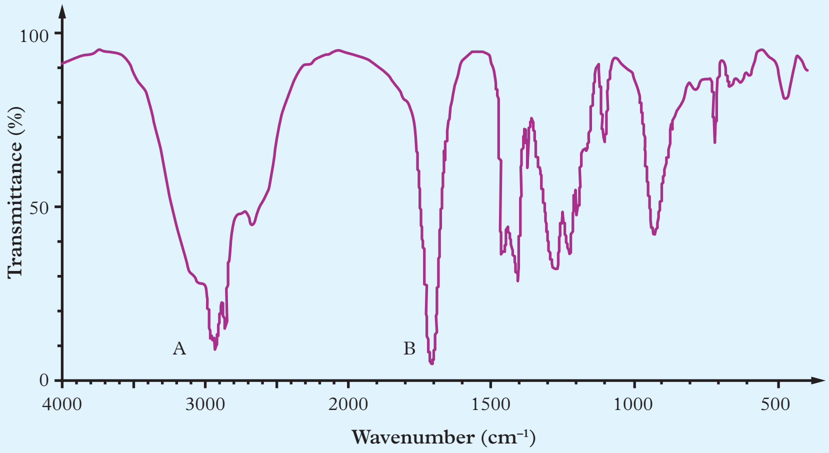
FIGURE 11 Some breathalysers use IR technology to detect ethanol in your breath.

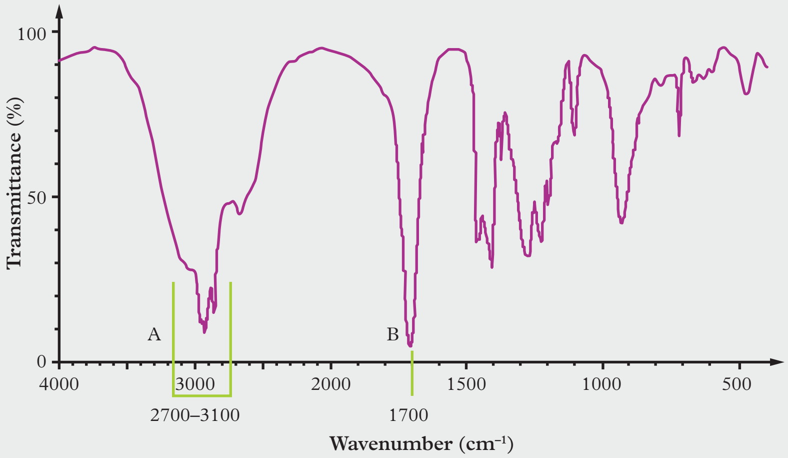
11.2 WORKED EXAMPLE

IDENTIFYING BONDS AND FUNCTIONAL GROUPS IN AN IR SPECTRUM

Identify the bonds responsible for the absorption peaks labelled A and B in the IR spectrum and provide a suggestion for the homologous series that the compound belongs to.



Solution

Think	Do
Step 1: Identify the wavenumber ranges that the tips of each peak occur at.	Peak A is at 2700–3100 cm^{-1} and peak B is at 1700 cm^{-1}. 
Step 2: Refer to your data book or Table 1 in this topic for characteristic IR ranges.	Peak A spans 2700–3100 cm^{-1}, indicating an O–H bond (it is not high enough to be N–H). This could be from an alcohol or a carboxylic acid. Peak B could be any of the C=O bonds listed around 1700 cm^{-1}. Pairing peak A and peak B together gives O–H and C=O; therefore, the compound is a carboxylic acid.

Study tip

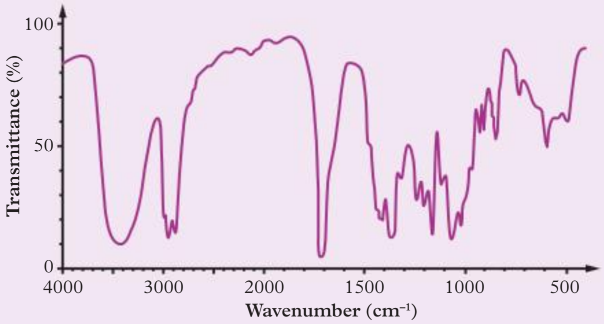
In some cases, spectra may not follow the common absorption patterns. *Always* be careful to double-check the wavenumbers, and be aware of the peak shapes.

11.2 CHECK YOUR LEARNING

Describe and explain

1 Explain why IR is useful to organic chemists.

Questions 2 and 3 refer to the IR spectrum shown below.



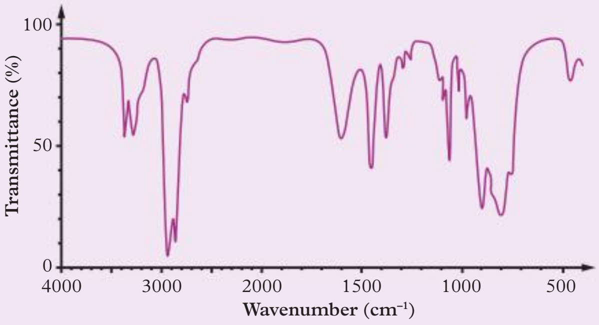
2 Describe the shape of the peaks (above the fingerprint region) in the IR spectrum.

3 Identify the bonds that each peak in the spectrum represents.

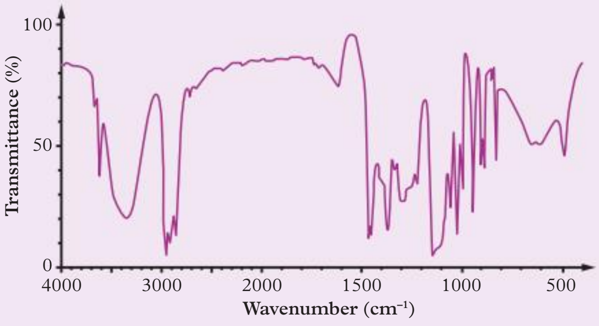
- **a** Peak at 1714 cm^{-1}
- **b** Peak at 2882 cm^{-1}
- **c** Peak at 3432 cm^{-1}

Apply, analyse and compare

4 Analyse the following spectrum and determine the functional group of the chemical that has produced the spectrum.

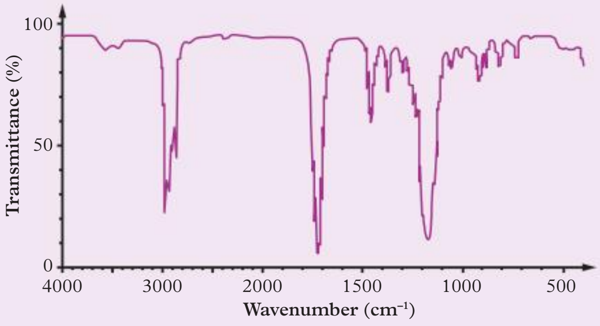


5 IR was performed on an alcohol to produce the following spectrum. Use the spectrum to explain why the compound is an alcohol.

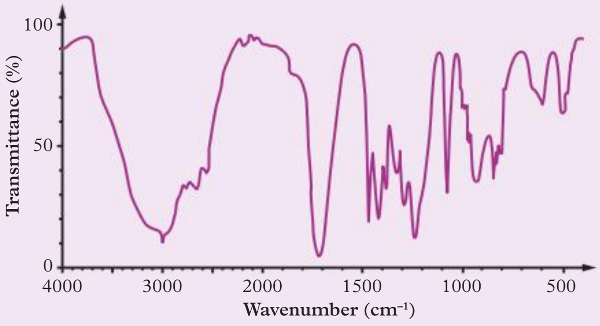


6 Two unknown organic compounds with the same molecular formula, $C_5H_{10}O_2$, are run through the IR spectrometer. They produce the following spectra:

Spectrum 1

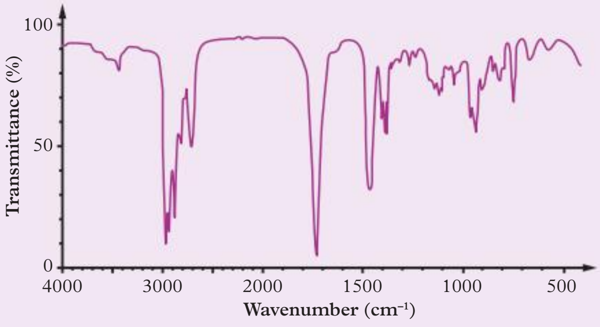


Spectrum 2



- **a** Analyse both spectra and determine the functional groups that may have produced them.
- **b** Draw the structural formula for all possible isomers of the compound that produced spectrum 1.
- **c** Name the compound that produced spectrum 2.

7 Analyse the following IR spectrum and list all of the functional groups that could not have produced it. Justify your answer.



11.3

Carbon-13 nuclear magnetic resonance spectroscopy

KEY IDEAS

In this topic, you will learn that:

- the number and type of carbon environments present in a compound can be determined by ^{13}C-NMR
- the number of peaks corresponds to the number of carbon environments in a particular compound
- chemical shift is used to identify the type of carbon environments present in a compound.

nuclear magnetic resonance (NMR)
an analytical technique in which the chemical environments in a compound can be determined by measuring the energy released from the nucleus of an atom after it absorbs electromagnetic radiation

electromagnetic radiation
the entire wavelength region on the electromagnetic spectrum that includes UV and visible light, infrared radiation, microwaves and radio waves

radio wave
a wavelength region of the electromagnetic spectrum above 10 cm

nuclear spin
the overall rotation of an atom's nucleus

^{13}C-NMR
carbon-13 nuclear magnetic resonance spectroscopy; the analysis of carbon-13 nuclear spin to identify carbon environments in a compound

Nuclear magnetic resonance (NMR) is a type of spectroscopy that can help determine the structure of complex organic compounds. IR uses infrared radiation to identify the bonds and functional groups within a compound, whereas NMR uses **electromagnetic radiation** in the **radio wave** region to determine how atoms are bonded together.

Nuclear magnetic resonance

NMR uses a magnetic field to induce **nuclear spin** in atoms with an odd number of protons and/or neutrons in their nucleus (e.g. ^{1}H and ^{13}C). Each proton and neutron has its own spinning motion. In atoms like ^{12}C, there is no overall nuclear spin because the subatomic particles spinning in opposite directions can cancel each other out. When there is an uneven number of protons and/or neutrons, there is an overall spin.

When a compound is placed in the NMR spectroscope, nuclei are at low energy spin states. Nuclei are then subjected to the electromagnetic field and enough energy is absorbed to cause the nucleus to flip into a high energy state (Figure 1).

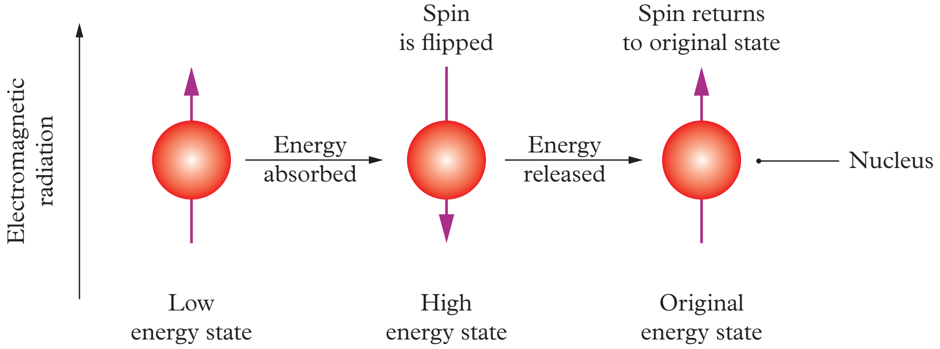


FIGURE 1 The change in energy to a nucleus caused by electromagnetic radiation

Eventually, the nucleus releases the energy it absorbs and flips back to its original state. The energy released (or the difference in energy between high and low spin states) is measured and used to generate an NMR spectrum.

NMR can be used to determine the number and nature of both ^{1}H and ^{13}C environments. While there are some principles that are consistent across both NMR types, there are many that differ. In this topic, we will look specifically at **^{13}C-NMR** and how it is used to identify carbon environments.

Carbon environments

chemical environment the atoms and bonds that surround a specific atom

carbon environment the atoms and bonds that surround a specific carbon atom

Atoms are considered to be in the same **chemical environment** (or chemically equivalent) if they absorb and release the same amount of energy (i.e. they will have the same spin). This is the case if they are covalently bonded to the same type of atoms and in the same way. Atoms in the same chemical environment will produce a single peak in an NMR spectrum. Chemically equivalent carbon atoms are considered to be in the same **carbon environment**. Let's look at some examples.

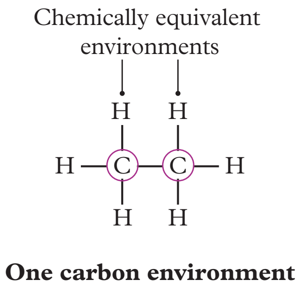


FIGURE 2 Carbon environments of ethane

The two carbon atoms in ethane are attached to the same number of hydrogen and carbon atoms (Figure 2). This means that they are chemically equivalent environments, i.e. ethane has only one carbon environment. The two carbon atoms in this environment are represented by the same single peak on the ^{13}C-NMR spectrum as both carbon nuclei absorb the same wavelength of radio waves to change spin states.

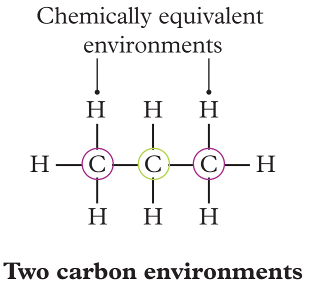


FIGURE 3 Carbon environments of propane

Propane has two carbon environments (Figure 3). One carbon environment is the carbon in CH_3 attached to a $-CH_2-$. This carbon environment occurs twice in the molecule but will only produce one peak on the ^{13}C-NMR spectrum. The second environment is the carbon in $-CH_2-$ that is attached to the two $-CH_3$ groups. This will form a second peak on the ^{13}C-NMR spectrum.

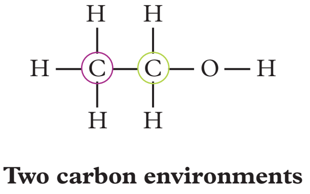


FIGURE 4 Carbon environments of ethanol

Ethanol also has two carbon environments: the carbon in $-CH_3$ and the carbon in $-CH_2$ (Figure 4). It will have two different peaks on its ^{13}C-NMR spectrum.

Molecules that are symmetrical will have fewer unique carbon environments and therefore, fewer peaks on their ^{13}C-NMR spectra.

Chemical environments at opposite ends of the molecule are only equivalent if there is symmetry in the molecule; i.e. the CH_3 groups at the ends of ethyl propanoate are not equivalent.

Study tip

When identifying carbon environments, look out for two things: what the carbon is attached to, and how these atoms are bonded to other atoms.

^{13}C-NMR chemical shift

chemical shift the amount of energy required to spin-flip a nucleus compared with TMS

The energy released when a nucleus spin-flips (or the energy required to change its spin state) is its **chemical shift**. Each different carbon will appear as a different peak at a different chemical shift on a ^{13}C-NMR spectrum. This is because different combinations of atoms will emit varying amounts of energy on return from spin-flip. As a result, different molecules will produce different NMR spectra.

A chemically neutral compound called tetramethylsilane (TMS) is used as a reference to keep the analysis of compounds consistent across all NMR spectrometers. This also allows chemists to share and compare results. TMS (Figure 5) is chemically inert and produces only one single peak. All chemical shift is compared against the shift of the TMS, which is 0 ppm.

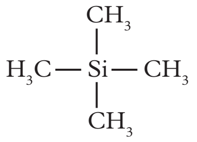


FIGURE 5 Tetramethylsilane (TMS) is used as a reference compound in NMR as it is chemically inert.

The shift is measured in parts per million ppm and represented by the symbol δ. The data book in your exam includes a table of ^{13}C-NMR chemical shift values, as shown in Table 1. TMS is also used in ^{1}H-NMR.

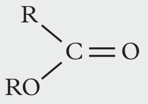
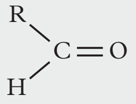
Analysing ^{13}C-NMR spectra

On a ^{13}C-NMR spectrum, chemical shift is shown along the *x*-axis. The number of peaks represents the number of different carbon environments in the molecule. There is no *y*-axis label and the height/area of the peaks is not proportional to the number of carbon atoms in each environment. The single peaks in ^{13}C-NMR do not tell you anything about the number of neighbouring carbon atoms either. This makes ^{13}C-NMR spectroscopy **low resolution**. Instead, the two main things you should focus on when analysing ^{13}C-NMR spectra are:

- the **number of peaks** – this will be the number of carbon environments in the molecules
- the **chemical shift** – when compared to shift data (Table 1), this will allow you to determine the carbon environment.

Let's look at the ^{13}C-NMR spectrum for propanoic acid as an example (Figure 6).

TABLE 1 Typical ^{13}C shift values relative to TMS = 0. These values can differ slightly in different solvents.

Type of carbon	Chemical shift (ppm)
$R{-}CH_3$	8–25
$R{-}CH_2{-}R$	20–45
$R_3{-}CH$	40–60
$R_4{-}C$	36–45
$R{-}CH_2{-}X$	15–80
$R_3C{-}NH_2$, $R_3C{-}NR$	35–70
$R{-}CH_2{-}OH$	50–90
$RC{\equiv}CR$	75–95
$R_2C{=}CR_2$	110–150
RCOOH	160–185
	165–175
	190–200
$R_2C{=}O$	205–220

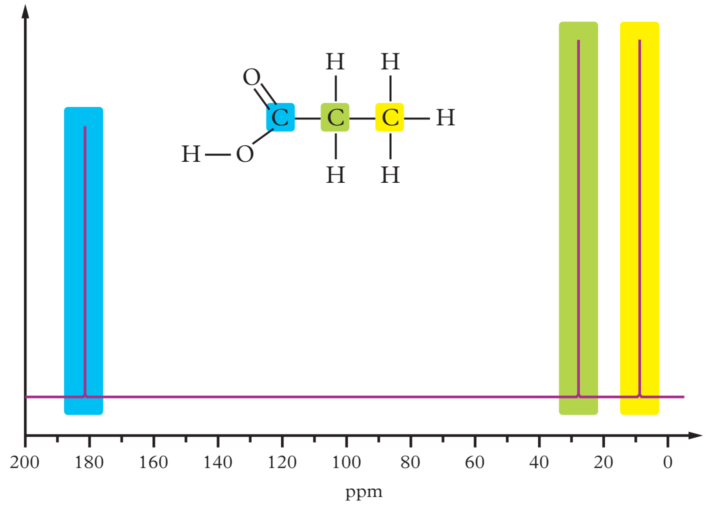


FIGURE 6 ^{13}C-NMR spectrum for propanoic acid

Propanoic acid has three different carbon environments: a $-CH_3$, a $-CH_2-$ and a $-COOH$. This corresponds to three peaks in the ^{13}C-NMR spectrum. We can look at the chemical shift for each carbon environment in your data book, or from Table 1.

- The first peak has a shift of 9 ppm, which corresponds to the shift for $-CH_3$, which is at 8–25 ppm.
- The second peak at 28 ppm corresponds to the shift for $-CH_2-$, which is at 20–45 ppm.
- The third peak at 181 ppm corresponds to the shift for $-COOH$, which is at 160–185 ppm.

You might have noticed that there is a lot of overlap in the shift values listed in your data book and in Table 1. This means analysing ^{13}C-NMR alone can be tricky. ^{13}C-NMR is best used in combination with IR spectra, MS spectra and ^{1}H-NMR spectra. However, when the names or structures of the possible compounds are given, it can be possible to identify the correct compound. See how to do this in Worked example 11.3 on the next page.

low resolution (in NMR spectroscopy) spectra that show only single peaks

Study tip

When analysing ^{13}C-NMR data, remember these two things: the number of peaks = the number of carbon environments, and the chemical shift values = the types of carbon environments.

Study tip

Chemical shift for ^{13}C-NMR overlaps a lot, so sometimes if you are considering a few options of compounds, it may be best to use the data book to look at what peaks are missing on the spectra and eliminate molecule options.

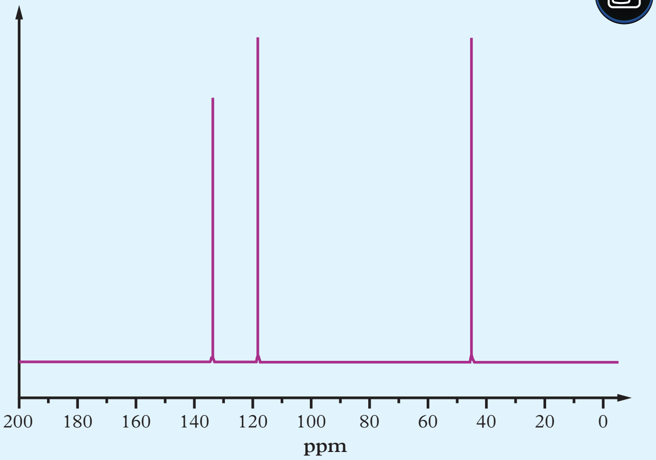
11.3 WORKED EXAMPLE

ANALYSING A ^{13}C-NMR SPECTRUM

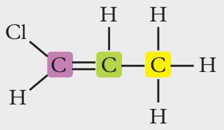
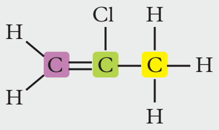
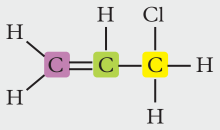
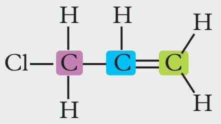
The ^{13}C-NMR spectrum shown belongs to one of three possible isomers of C_3H_5Cl:

- 1-chloroprop-1-ene
- 2-chloroprop-1-ene
- 3-chloroprop-1-ene.

Analyse the ^{13}C-NMR spectrum and identify which of the isomers it belongs to.



Solution

<table>
<tr><th>Think</th><th>Do</th></tr>
<tr><td>Step 1: Draw the structures of the isomers so you can determine the number and type of carbon environments in each.</td><td>
1-chloroprop-1-ene
Carbon environments:
CHCl CH– –CH_3

2-chloroprop-1-ene
Carbon environments:
CH_2 CCl– –CH_3

3-chloroprop-1-ene
Carbon environments:
CH_2 CH –CH_2Cl
The spectrum shows three peaks, and each isomer has three carbon environments. None of the isomers can be ruled out at this stage.</td></tr>
<tr><td>Step 2: Consider the chemical shifts from the spectrum.</td><td>From right to left, the chemical shifts for each peak are:
Peak 1: 45 ppm
Peak 2: 118 ppm
Peak 3: 134 ppm
Looking at the ^{13}C-NMR shift in the data book or Table 1, there is no peak at 8–25 ppm, meaning there is no –CH_3 group in the compound. This means that the compound cannot be 1-chloroprop-1-ene or 2-chloroprop-1-ene, as both have –CH_3 groups. Therefore, this is the spectrum for 3-chloroprop-1-ene.</td></tr>
<tr><td>Step 3: Check the other shift against what is given in the data book to confirm.</td><td>
<table><tr><th></th><th>Type of carbon</th><th>Chemical shift (ppm)</th></tr><tr><td>Peak at 45 ppm →</td><td>R — CH_2— X</td><td>15–80</td></tr><tr><td>Two peaks at 118 ppm and 134 ppm →</td><td>$R_2C{=}CR_2$</td><td>110–150</td></tr></table>
The peak at 45 ppm is the CH_2–Cl and the two others are within the range of double-bonded carbon atoms, confirming it is the spectrum for 3-chloroprop-1-ene.</td></tr>
</table>

11.3 CHECK YOUR LEARNING

Describe and explain

1 Explain what you can determine using ^{13}C-NMR.

2 A compound with the formula C_4H_8O was analysed using ^{13}C-NMR.

  a Draw the unbranched ketone and aldehyde with this molecular formula.

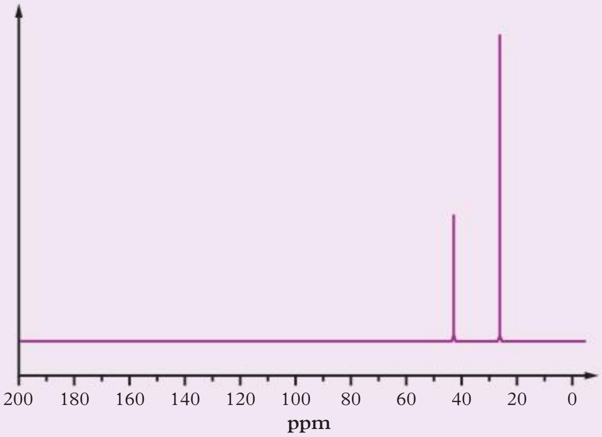
  b Describe the peaks you would get on a ^{13}C-NMR spectrum of each compound.

  c Explain if you could determine if it was a ketone or an aldehyde by using just ^{13}C-NMR.

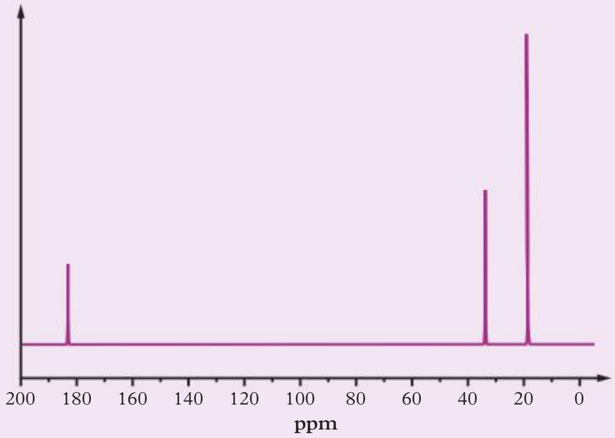
3 Describe the peaks you would expect to see in the ^{13}C-NMR spectrum for pentan-2,4-diol.

Apply, analyse and compare

4 A compound containing carbon, hydrogen and nitrogen produces the ^{13}C-NMR spectrum shown. Discuss and draw the structures of all the possible compounds that the ^{13}C-NMR spectrum could represent.



5 The following ^{13}C-NMR spectrum is produced by an organic compound with the molecular formula $C_4H_8O_2$.



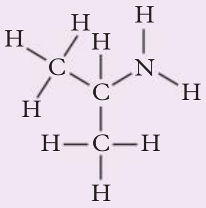
  a Explain why there are only three carbon environments, even though there are four carbon atoms in the molecule.

  b Suggest a molecule that could produce this ^{13}C-NMR spectrum.

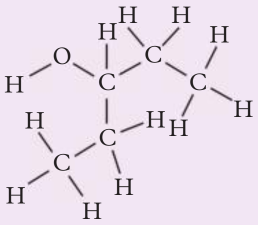
Design and discuss

6 Four organic compounds are shown below. Discuss how many carbon environments are in each.

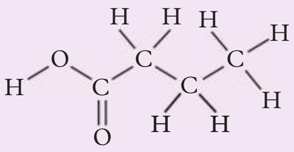
a



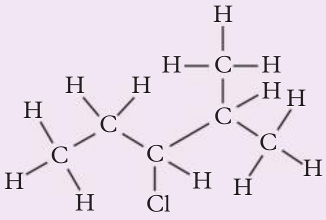
b



c



d



11.4 Hydrogen-1 nuclear magnetic resonance spectroscopy

KEY IDEAS

In this topic, you will learn that:

- the number and type of hydrogen environments present in a compound can be determined by ^{1}H-NMR
- peak area is used to determine the number of hydrogen atoms in an environment in a particular compound
- chemical shift is used to identify the type of hydrogen environments present in a compound
- splitting patterns in ^{1}H-NMR are used to identify the number of neighbouring hydrogen atoms in each environment using the $n + 1$ rule.

^{1}H-NMR
hydrogen-1 (or proton) nuclear magnetic resonance spectroscopy; the analysis of hydrogen-1 nuclear spin to identify hydrogen environments in a compound

hydrogen environment
the atoms and bonds that surround a specific hydrogen atom

Study tip

Hydrogen-1 NMR is also known as proton NMR, because a hydrogen nucleus is just a proton.

Hydrogen-1 (or proton) NMR uses the same principles as ^{13}C-NMR: a magnetic field is used to induce spin in the odd-numbered nucleus of a hydrogen-1 atom, creating a spectrum similar to a ^{13}C-NMR spectrum. The analysis of a **^{1}H-NMR** spectrum gives more information about an organic compound than that of ^{13}C-NMR.

Hydrogen environments

Just as ^{13}C-NMR is used to identify carbon environments, the same can be done with ^{1}H-NMR and **hydrogen environments** in different compounds.

In Topic 11.3, we looked at ethane, propane and ethanol and determined the number of carbon environments in each molecule. Let's look at the same three molecules to examine the number of hydrogen environments in each.

Previously, we determined that the two carbon atoms in ethane were chemically equivalent. Ethane is symmetrical and therefore, has one carbon environment. We can also see that all six of the hydrogen atoms in ethane are in the same environment: each is part of a $-CH_3$ group that is attached to another $-CH_3$ group. The hydrogen environments are chemically equivalent and will produce a single peak on a ^{1}H-NMR spectrum (Figure 1).

Propane, which has two carbon environments, also has two hydrogen environments. One hydrogen environment consists of the hydrogen atoms in CH_3 attached to a CH_2. This environment occurs twice in the molecule, so any hydrogen atoms in the CH_3 environment are chemically equivalent and will produce a single peak in the ^{1}H-NMR spectrum (Figure 2). The other hydrogen environment consists of the hydrogen atoms in $-CH_2-$ attached to the two $-CH_3$ groups. This will form the second peak on the ^{1}H-NMR spectrum.

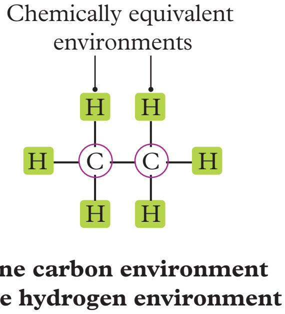


FIGURE 1 The carbon and hydrogen environments of ethane

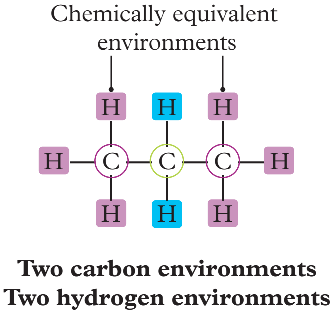


FIGURE 2 The carbon and hydrogen environments of propane

Ethanol has two carbon environments but three hydrogen environments (Figure 3). It has a $-CH_3$ connected to a $-CH_2-$ that will produce a peak. The hydrogen atoms in $-CH_2-$ will produce a different second peak, and the final hydrogen atom in –OH will produce the third peak in its ^{1}H-NMR spectrum.

Molecules that are symmetrical will have fewer unique hydrogen environments and therefore, fewer peaks in their ^{1}H-NMR spectra. Each different hydrogen will appear as a different peak at a different chemical shift on the ^{1}H-NMR spectrum, just like in ^{13}C-NMR spectroscopy.

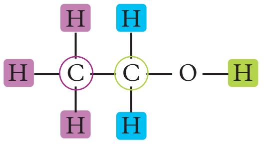


FIGURE 3 The carbon and hydrogen environments of ethanol

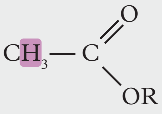
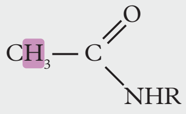
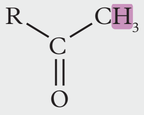
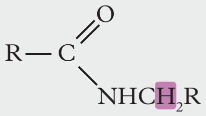
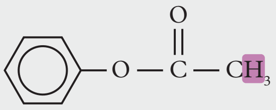
Study tip

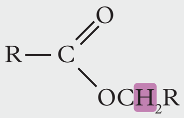
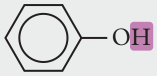
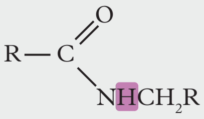
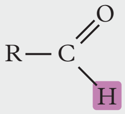
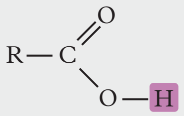
When identifying hydrogen environments, look out for two things: what the hydrogen is attached to, and how these atoms are bonded to other atoms.

^{1}H-NMR chemical shift

^{1}H-NMR also uses TMS as a reference compound for chemical shift, measured in ppm. However, in ^{1}H-NMR, the chemical shift is a lot lower than in ^{13}C-NMR. Table 1 shows chemical shifts for different protons in ^{1}H-NMR, highlighted in purple. This is also given in your data book.

TABLE 1 Typical ^{1}H-NMR shift values relative to TMS = 0. These values can differ slightly in different solvents.

Type of proton	Chemical shift (ppm)
$R-CH_3$	0.9–1.0
$R-CH_2-R$	1.3–1.4
$RCH=CH-CH_3$	1.6–1.9
R_3-CH	1.5
 or 	2.0
	2.1–2.7
$R-CH_2-X$ (X = F, Cl, Br or I)	3.0–4.5
$R-CH_2-OH$, $R_2-CH-OH$	3.3–4.5
	3.2
$R-O-CH_3$ or $R-O-CH_2R$	3.3–3.7
	2.3

Type of proton	Chemical shift (ppm)
	3.7–4.8
$R-O-H$	1–6
$R-NH_2$	1–5
$RHC=CHR$	4.5–7.0
	4.0–12.0
	6.9–9.0
	8.1
	9.4–10.0
	9.0–13.0

Analysing ^{1}H-NMR spectra

^{1}H-NMR spectra look similar to ^{13}C-NMR, with chemical shift shown in the *x*-axis and no *y*-axis label. The number of peaks represents the number of different hydrogen environments. A key difference is that ^{1}H-NMR can be **high resolution**.

high resolution
(in NMR spectroscopy) spectra that show finer details for each peak, such as peak splitting

Study tip

Chemical shift values for ^{1}H-NMR are included in your data book, so you do not have to remember them. Practise using your data book so that you are familiar with the process for when you sit the exam.

Like ^{13}C-NMR, low resolution ^{1}H-NMR is much simpler and shows only single peaks (Figure 4a). On the other hand, high resolution ^{1}H-NMR includes further details about each peak (Figure 4b).

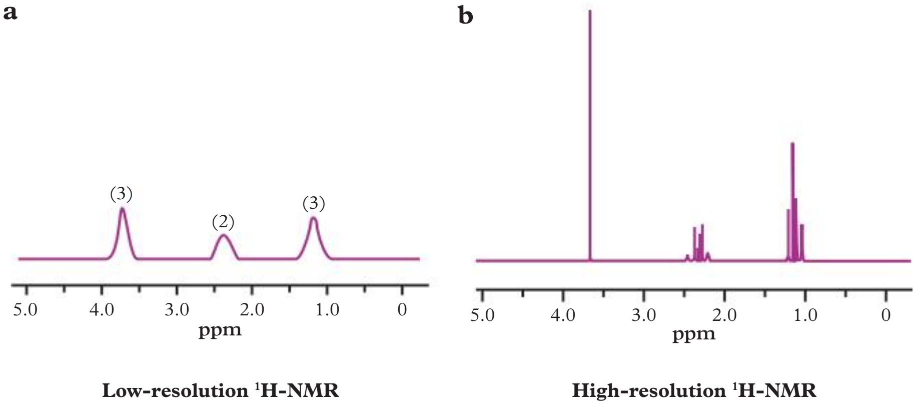


FIGURE 4 **a** Low-resolution and **b** high-resolution ^{1}H-NMR spectra of methyl propanoate

There are three main things to consider when analysing ^{1}H-NMR spectra:

- the **number of peaks** – this will be the number of hydrogen environments in the molecule
- the **chemical shift** – when compared to shift data (Table 1), this will allow you to determine the hydrogen environment
- the **peak area** – this will give you the number of hydrogen atoms in that environment.

When you are analysing high-resolution ^{1}H-NMR spectra, there is one more thing to consider:

- **peak splitting** – this will provide information about neighbouring hydrogen atoms, using the $n + 1$ rule.

peak area
the area under a peak

integration curve
a curve in which the height corresponds to the relative peak area

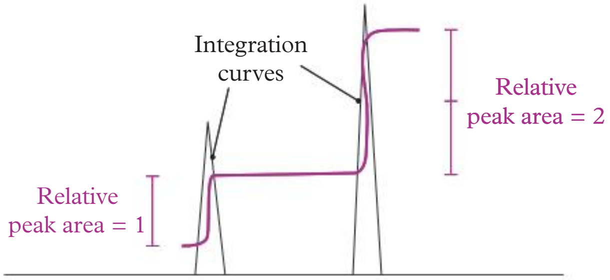


FIGURE 5 The relative heights of the integration curves can be used to determine relative peak area.

Peak area

Unlike in ^{13}C-NMR, **peak area** is very useful in ^{1}H-NMR. It is proportional to the number of hydrogen atoms in a hydrogen environment. In most of the spectra you will encounter, it is too difficult to measure the peak area. Luckily, we can look at **integration curves**, which represent the area under each peak.

You can measure and compare the heights of the integration curves in a ^{1}H-NMR spectrum to determine the ratio of hydrogen atoms in each environment (Figure 5).

The integration curve for the peak on the right has double the height of the integration curve for the peak on the left. Therefore, the hydrogen environment represented by the right peak must have double the number of hydrogen atoms compared to the left peak. Because the peak areas are relative, you cannot use them to determine the total number of hydrogen atoms in each environment, unless you know the molecular formula.

You may be asked to measure the heights of these curves in NMR spectra, or they may be given to you in a table or above each peak. Worked example 11.4A takes you through how to use integration curves from a low resolution ^{1}H-NMR spectrum to determine the number of hydrogen atoms in each environment.

11.4A Worked example
Find me in your obook pro

11.4A Worked example
Video demonstration

Peak splitting and the *n* + 1 rule

Peak splitting in ^{1}H-NMR is caused by neighbouring hydrogen atoms. There are some rules to consider when looking at peak splitting:

- An environment is 'neighbouring' if it is up to three bonds away.
- To calculate the number of finer peaks caused by splitting, use the ***n* + 1 rule**, where *n* = the number of neighbouring hydrogen atoms.
- A hydrogen in –O–H will never cause splitting and will not be split.
- Equivalent hydrogen environments will not cause each other to split; for example, ethane will have a single peak.
- If a hydrogen is surrounded by two equivalent peaks, its splitting pattern will be determined by all of its surrounding hydrogen atoms; for example, a propane ^{1}H-NMR will be a triplet (3) and a septet (7).
- A hydrogen with two different (non-equivalent) environments on either side will undergo complex coupling or splitting. This is not covered in VCE Chemistry.

The rules for naming peaks are shown in Table 2.

peak splitting
the further splitting of individual peaks caused by neighbouring hydrogen atoms

***n* + 1 rule**
a rule stating that hydrogen NMR peak will split into finer peaks; the number of finer peaks = the number of neighbouring hydrogen atoms + 1

TABLE 2 Fine peak splitting nomenclature

Number of fine peaks	Name	Neighbouring hydrogen atoms
1	Singlet	No neighbouring hydrogen atoms
2	Doublet	Next to a –CH–
3	Triplet	Next to a –CH_2–
4	Quartet	Next to a –CH_3

If splitting results in five or more fine peaks, it can be referred to as a **multiplet**, or by common prefixes (quintet, sextet, etc.). You will often be given the number of fine peaks in brackets along with the name; e.g. 'doublet (2)'.

multiplet
a hydrogen NMR peak that has split into five or more fine peaks

Let's apply what we have learnt by analysing the high resolution ^{1}H-NMR spectrum of ethanol (Figure 6). There are three main peaks and each of the three peaks is split into multiple finer peaks.

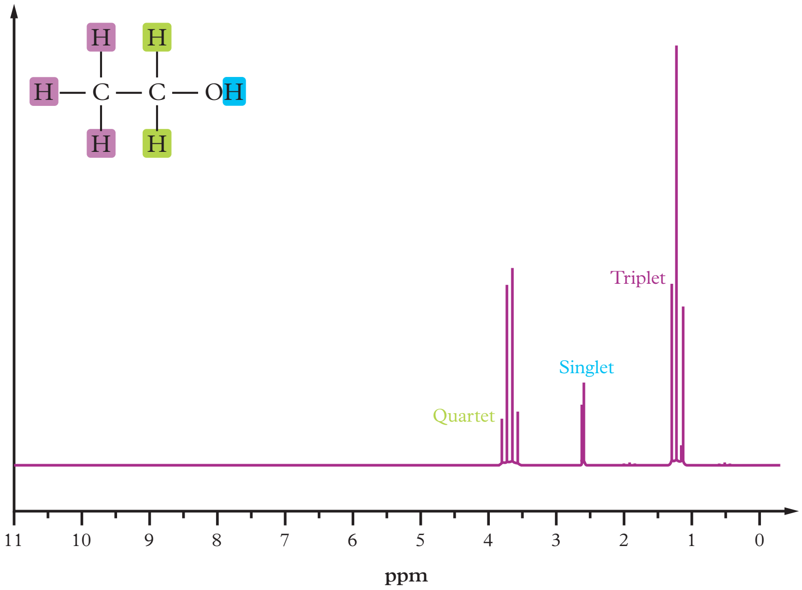
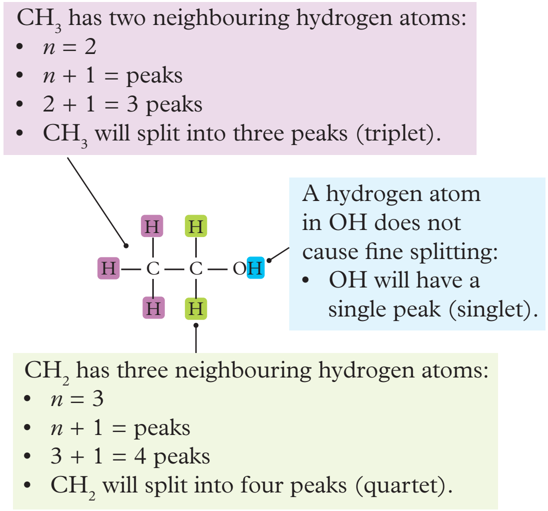


FIGURE 6 The ^{1}H-NMR spectrum for ethanol

The number of fine peaks for ethanol shown in Figure 6 on the previous page are calculated below in Figure 7.



CH_3 has two neighbouring hydrogen atoms:

- $n = 2$
- $n + 1$ = peaks
- $2 + 1 = 3$ peaks
- CH_3 will split into three peaks (triplet).

A hydrogen atom in OH does not cause fine splitting:

- OH will have a single peak (singlet).

CH_2 has three neighbouring hydrogen atoms:

- $n = 3$
- $n + 1$ = peaks
- $3 + 1 = 4$ peaks
- CH_2 will split into four peaks (quartet).

FIGURE 7 Fine peak splitting calculations for ethanol

Remember that when you analyse low resolution ^{1}H-NMR, you need to consider the number of peaks, peak area and chemical shift. When analysing high resolution ^{1}H-NMR, you must also consider peak splitting. Peak splitting gives you the neighbouring hydrogen environments and therefore, more information about the full structure of an organic molecule.

Figure 8 shows a high resolution ^{1}H-NMR spectrum of chloroethane. To analyse it, let's consider the number of peaks, peak area, peak splitting, and chemical shift.

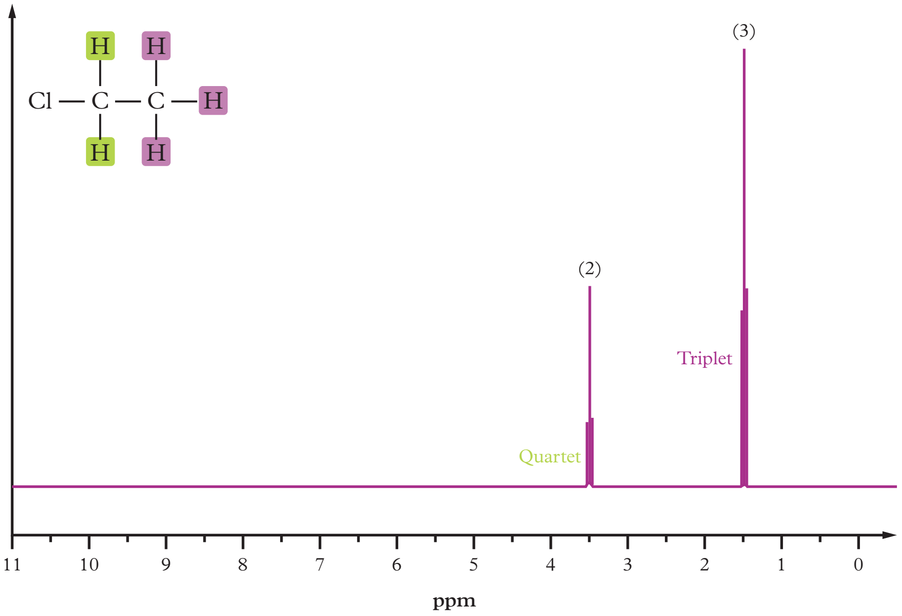
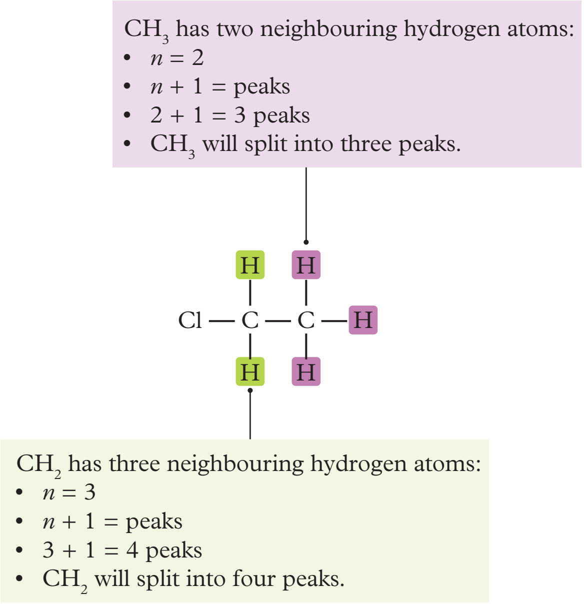


FIGURE 8 The ^{1}H-NMR spectrum of chloroethane

- Number of peaks: Chloroethane has two hydrogen environments, the $-CH_3$ (in purple) and the $-CH_2Cl$ (in green). This gives two peaks in the ^{1}H-NMR spectrum.
- Peak area: In this spectrum, the peak areas have been provided in brackets above the peaks. This is the ratio of hydrogen atoms. The (2) and (3) tell us that the ratio of hydrogen atoms in chloroethane is 2 : 3. In this case, the (2) and (3) correspond to the $-CH_2$ and $-CH_3$ groups, respectively.

- Peak splitting: The –CH_3 peak has been split into three finer peaks (a triplet), and the –CH_2Cl has split into four fine peaks (a quartet). The calculations of finer peaks for both –CH_3 and –CH_2Cl are shown in Figure 9.



CH_3 has two neighbouring hydrogen atoms:
- $n = 2$
- $n + 1$ = peaks
- $2 + 1 = 3$ peaks
- CH_3 will split into three peaks.

CH_2 has three neighbouring hydrogen atoms:
- $n = 3$
- $n + 1$ = peaks
- $3 + 1 = 4$ peaks
- CH_2 will split into four peaks.

FIGURE 9 Fine peak splitting calculations for chloroethane

- Chemical shift: Because so many combinations of atoms exist, you might not always find an exact chemical shift from the data book that matches the shift on the spectrum. In Figure 7, the first chemical shift in the ^{1}H-NMR spectrum of chloropropane is at 1.4 ppm. This does not match any of the shift data, as R–CH_3 is listed as 0.9–1.0 ppm. This is okay, as we already know it is a –CH_3 from the peak area and splitting patterns. The next peak in the spectrum is at 3.5 ppm, which corresponds to the chemical shift in the data book for R–CH_2–X (X = F, Cl, Br, or I). We can confirm that the second peak is CH_2Cl from the chemical shift.

Study tip

When analysing ^{1}H-NMR, leave the chemical shift data until last. Sometimes, the data book will not have the range for the H environment you are looking for, or the shift differs slightly because of different solvents used to analyse the samples.

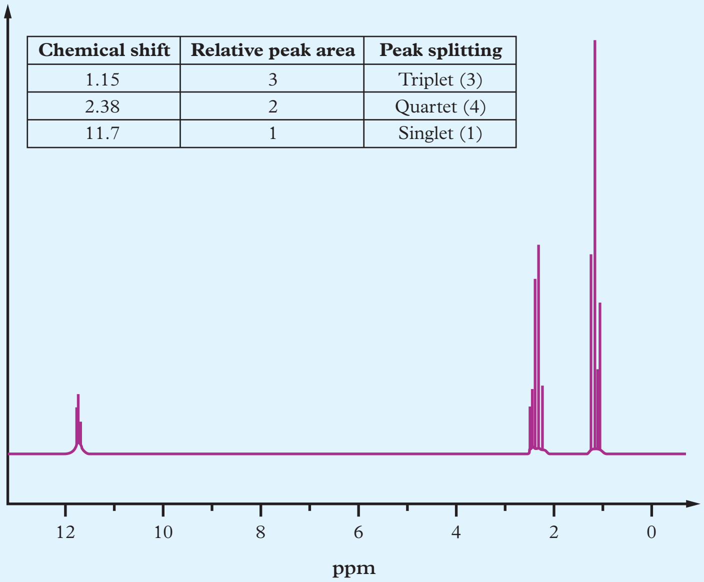
FIGURE 10 Samples for NMR are loaded into small glass tubes.

FIGURE 11 Computer software can be used to help analyse ^{1}H-NMR data.

11.4B WORKED EXAMPLE

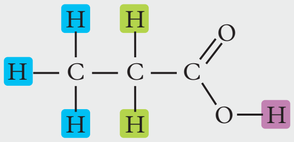
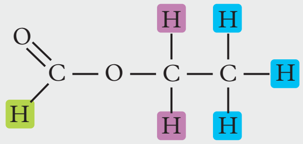
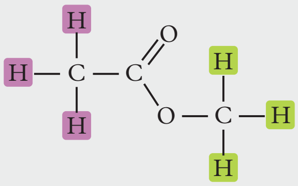
ANALYSING A ^{1}H-NMR SPECTRUM

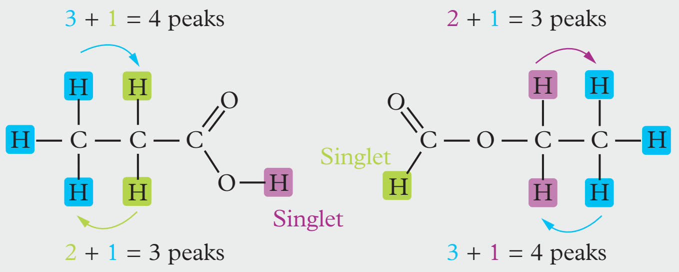
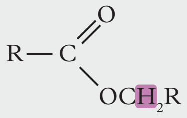
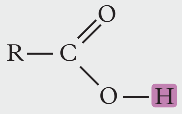
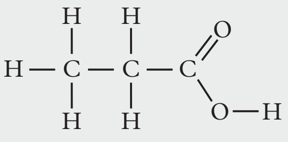
An organic compound with the molecular formula $C_3H_6O_2$ has the ^{1}H-NMR spectrum shown. Determine the structural formula and correct IUPAC name for the compound.



Chemical shift	Relative peak area	Peak splitting
1.15	3	Triplet (3)
2.38	2	Quartet (4)
11.7	1	Singlet (1)

Solution

Think	Do
Step 1: Consider the isomers that could exist for $C_3H_6O_2$.	It could be a carboxylic acid or an ester – i.e., any of propanoic acid, methyl ethanoate or ethyl methanoate.
Step 2: Draw the structural formulas and identify the number of peaks that each of the isomers could have on a ^{1}H-NMR spectrum.	3 peaks  **Propanoic acid** 3 peaks  **Ethyl methanoate** 2 peaks  **Methyl ethanoate**

<table>
<tr><th>Think</th><th>Do</th></tr>
<tr><td>Step 3: Determine whether you can rule any of the isomers out based on the number of peaks and the peak area given in the ^{1}H-NMR spectrum.</td><td>The ^{1}H-NMR spectrum has three main peaks, which rules out methyl ethanoate. That leaves us with propanoic acid and ethyl methanoate.
Peak area is 3 : 2 : 1, but both propanoic acid and ethyl methanoate have hydrogen environments with 3, 2 and 1 hydrogen atoms, so this does not narrow it down.</td></tr>
<tr><td>Step 4: Determine whether you can rule any of the isomers out based on peak splitting.</td><td>The peak splitting patterns in the spectrum includes a triplet, quartet and a singlet.

Propanoic acid Ethyl methanoate
Both propanoic acid and ethyl methanoate will produce a triplet, quartet, and a singlet. Neither isomer can be ruled out based on peak splitting.</td></tr>
<tr><td>Step 5: Determine whether you can rule any of the isomers out based on chemical shift.</td><td>Chemical shift is 1.15, 2.38 and 11.7 ppm. The shift data from the data book identifies two different protons that are unique to each of the isomers we are considering.
<table>
<tr><th>Type of proton</th><th>Chemical shift (ppm)</th></tr>
<tr><td> in ethyl methanoate</td><td>3.7–4.8</td></tr>
<tr><td> in propanoic acid</td><td>9.0–13.0</td></tr>
</table>
There is no peak at 3.7–4.8 ppm in the spectrum, so it cannot be ethyl methanoate. There is a peak at 11.7 ppm, so this will correspond to a hydrogen in a –COOH. This confirms that the compound is propanoic acid.</td></tr>
<tr><td>Step 6: Draw the full structural formula of the compound (showing all bonds) and name it using the IUPAC rules.</td><td>
Propanoic acid</td></tr>
</table>

If you are feeling confident, practise identifying compounds using their ^{1}H-NMR spectra in Challenge 11.4.

11.4 Challenge
Find me in your obook pro

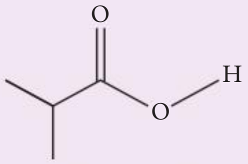
11.4 CHECK YOUR LEARNING

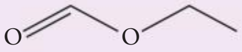
Describe and explain

1 Explain what you can determine using ^{1}H-NMR.

2 Describe the peaks, chemical shift and splitting you would expect to see in the ^{1}H-NMR spectra of the compounds shown:

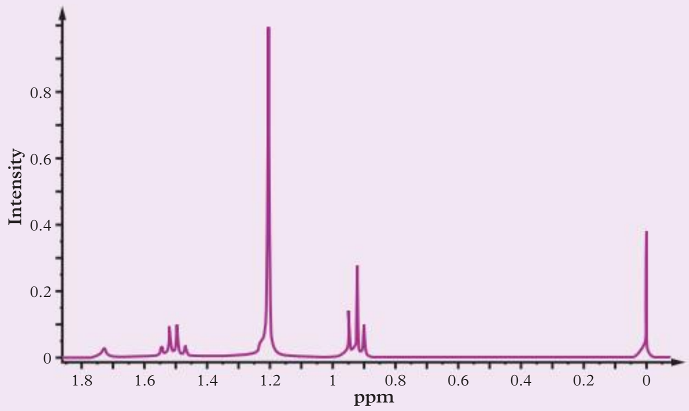
a 

b 

c 

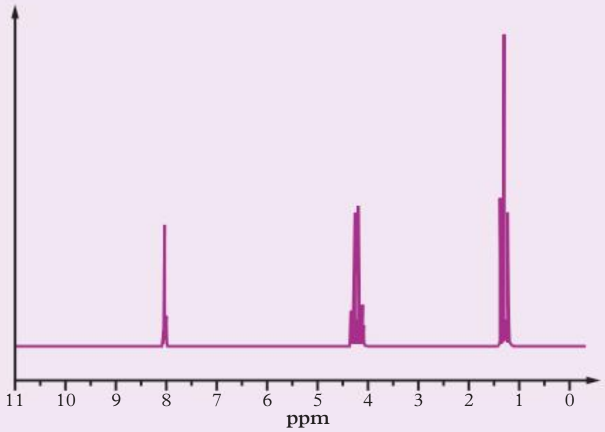
3 Explain the $n + 1$ rule for ^{1}H-NMR.

4 Describe the splitting pattern in the ^{1}H-NMR spectrum below.



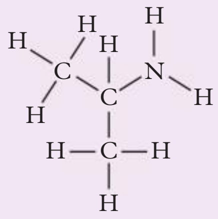
Apply, analyse and compare

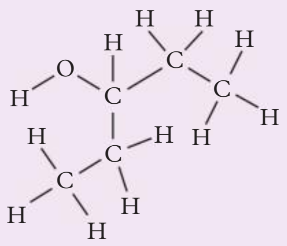
5 Analyse the following ^{1}H-NMR spectrum and explain why it is the spectrum for ethyl methanoate.

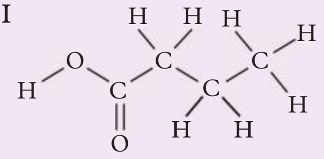


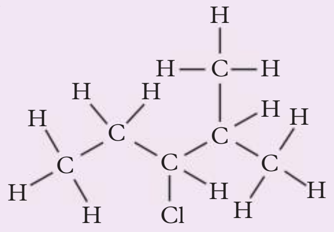
Design and discuss

6 Four organic compounds are shown below:

I 

II 

III 

IV 

a Discuss how many hydrogen environments there are in each.

b Describe the peak area ratio you would expect to see for the ^{1}H-NMR spectrum of each molecule.

11.5 Determining structures of simple organic compounds

KEY IDEAS

In this topic, you will learn that:

- you can use MS, IR, ^{13}C-NMR and ^{1}H-NMR to identify unknown organic compounds.

Over the course of this chapter, you have learnt about techniques including MS, IR spectroscopy, and ^{1}H-NMR and ^{13}C-NMR spectroscopy. In this topic, you will see that each of these techniques provides a different, useful piece of information and that, when put together, can help determine the structure of unknown organic compounds.

Combining spectroscopy techniques

Each type of instrumental analysis gives useful information about a different part of the structure of an unknown organic compound. Figure 1 summarises what each spectroscopy technique can be used to determine.

Mass spectrometry	Infrared spectrometry	^{13}C-NMR	^{1}H-NMR
• Molar mass • Individual fragments (sometimes)	• Functional groups present	• Number and type of carbon environments present	• Number and type of hydrogen environments present • Ratio of hydrogen atoms in each environment • Number of hydrogen atoms in neighbouring environments

FIGURE 1 Information gained from each spectroscopy technique

There is no exact order in which you must look at the different spectra. In fact, when you are completing combination-type questions, you may find yourself flipping between the different spectra constantly. Each time you identify something, remember to check that it works with all the information you are provided.

Worked examples 11.5A and 11.5B walk you through the process of using multiple sets of information to identify a compound. Lots of information can be presented in combination-type questions, so make sure to practise interpreting the different spectra using the practice questions.

Once you have gained some confidence, give Challenge 11.5 a go.

11.5
Challenge
Find me in your obook pro

11.5A
PRACTICAL: CASE STUDY

How can analytical techniques be used to differentiate between isomers?
Go to page **522**.

11.5B
PRACTICAL: CASE STUDY

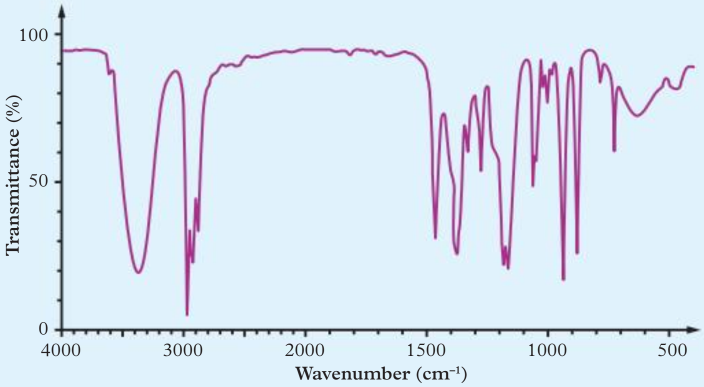
Lab clean-up duty – what's in the bottle?
Go to your **Student obook pro**.

11.5A WORKED EXAMPLE

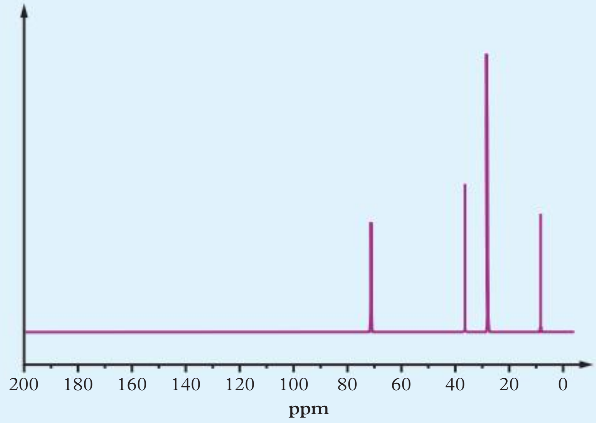
IDENTIFYING AN UNKNOWN COMPOUND USING A COMBINATION OF SPECTROSCOPY TECHNIQUES

An unknown compound containing carbon, hydrogen and oxygen has a molar mass of 88.0 g mol^{-1}. Using the three spectra shown, determine the name and structure (showing all bonds) of the unknown compound.

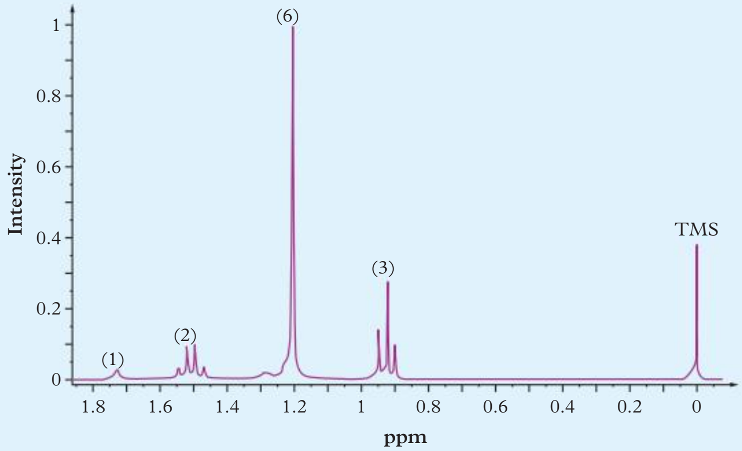
IR spectrum:



^{13}C-NMR spectrum:



^{1}H-NMR spectrum (relative peak area is indicated in brackets):



11.5B Worked example
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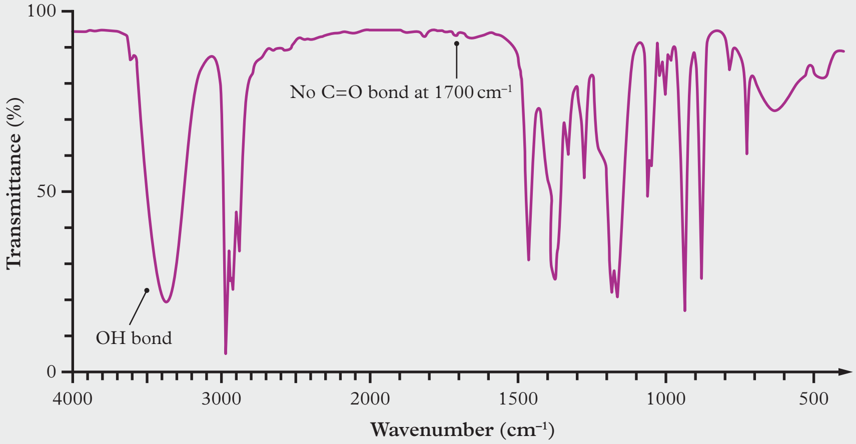
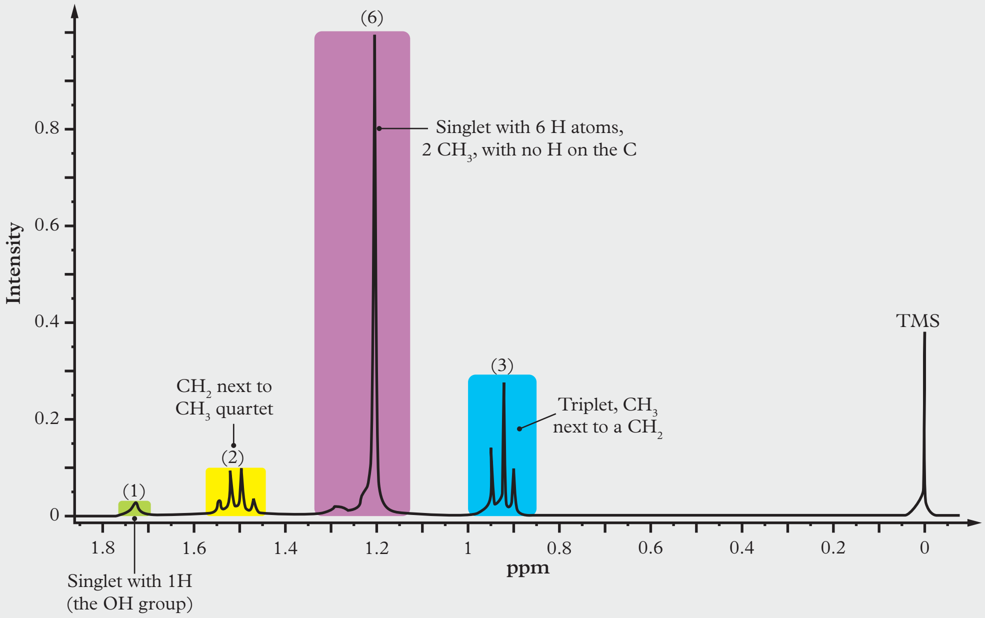
11.5B Worked example
Video demonstration

Study tip

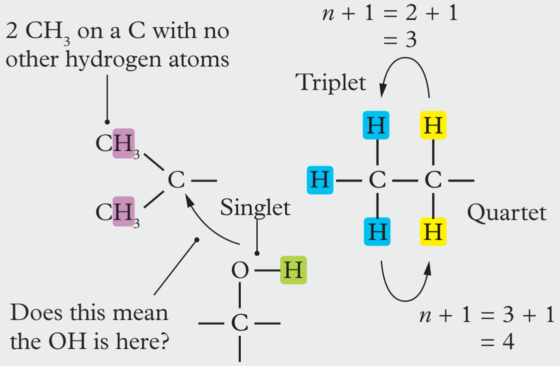
'Practice makes perfect' is something you might hear all the time, but it is certainly true for combining spectroscopy techniques. Since there are no set rules to follow, you may have to break down every question differently. The more you practise, the better you will get at this.

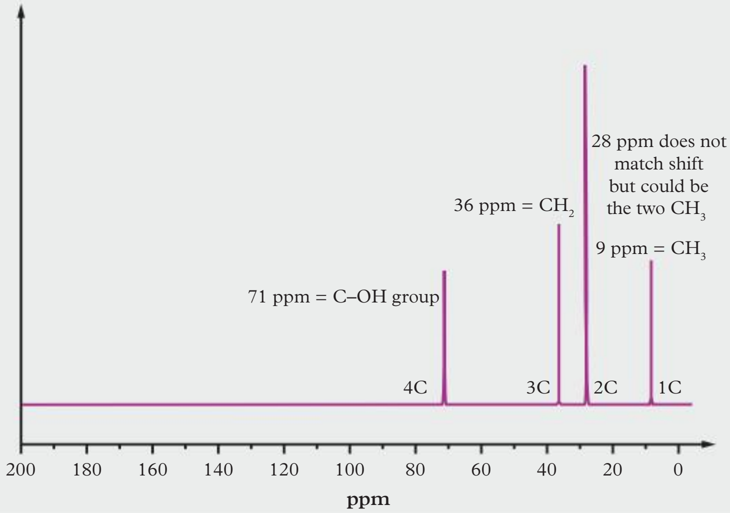
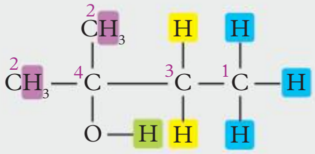
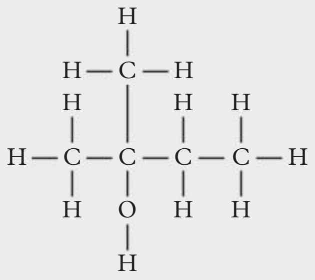
Solution

Think	Do
Step 1: Look at each spectrum and the information in the question stem to get a clear view of what you may have.	There is carbon, hydrogen and oxygen present, and the compound has a molar mass of 88 g mol^{-1}. The presence of carbon, hydrogen and oxygen gives us a few options. The compound could be an alcohol, ester, ketone, aldehyde or carboxylic acid. We will need to look at the IR to see what functional groups we have.
Step 2: Examine the IR spectrum to determine the functional groups present.	Looking at the IR, we can tell that the compound has a hydroxyl functional group (–OH) from the very strong, broad peak in the 3200–3600 cm^{-1} region. There is no strong, narrow peak in the 1700 cm^{-1} region, so the compound cannot be a carboxylic acid, aldehyde, ketone or ester. It must be an alcohol. 
Step 3: Use the IR spectrum and the molar mass to determine the molecular formula of the compound.	$M = 88.0\ \text{g mol}^{-1}$ Remember the general formula for an alcohol is $C_nH_{2n+2}O$. $M(\text{C}) = 12.0$, $M(\text{H}) = 1.0$ and $M(\text{O}) = 16.0$ $(n \times 12.0) + ((2n + 2) \times 1.0) + 16.0 = 88.0$ $12n + 2n + 2 + 16.0 = 88.0$ $14n = 70$ $n = 5$ So, the molecular formula is $C_5H_{12}O$.
Step 4: Look at the ^{13}C-NMR and ^{1}H-NMR spectra for carbon and hydrogen environments, and determine if any isomers with this molecular formula can be ruled out.	^{13}C-NMR: four carbon environments ^{1}H-NMR: four hydrogen environments Pentanol can be ruled out, as it would have five carbon and five hydrogen environments. The fact that the compound has four H and C environments indicates that there are two sets of carbon atoms and hydrogen atoms that have the same environment. This means that the compound is likely to contain two chemically equivalent methyl groups.

Think	Do
Step 5: Check the ^{1}H-NMR spectrum for more information and lay out all your information so that you can see what you have. You could annotate the spectrum, make a table, or use dot points. Colour-coding with different highlighters may also help. After this, you can sketch out options based on what you have so far.	

Chemical shift (ppm)	Relative peak area	Peak splitting	Indicates
1.7	1	Singlet	–OH
1.5	2	Quartet	CH_2 next to a CH_3
1.2	6	Singlet	2 × CH_3 on a carbon atom, with no other hydrogen atoms
0.9	3	Triplet	CH_3 next to a CH_2



Think	Do
Step 6: Look at the ^{13}C-NMR spectrum to confirm carbon environments.	The shift for the second peak does not match anything in the data book, and the shift at 71 ppm from the data book says CH_2OH. Since we cannot have CH_2OH, we can make an educated guess that this is our C–O–H peak. 
Step 7: Put everything together to determine the final structure and recheck this structure against all the information provided.	The information we have is $M(C_5H_{12}O)$ = 88.0 g mol^{-1}, with four H environments (highlighted) and four C environments (indicated by the purple numbers). It gives us this structure. 
Step 8: Draw the full structural formula, showing all bonds, and name the compound.	 2-methylbutan-2-ol

11.5 CHECK YOUR LEARNING

Apply, analyse and compare

1 Consider the molecule propan-1-amide.

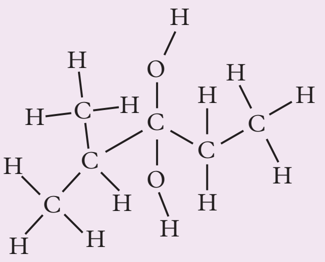
  a Identify the *m*/*z* of the molecular ion peak in the mass spectrum of this compound.

  b Describe the IR spectrum that you would obtain for this compound.

  c Explain how many peaks you would see on a low resolution ^{1}H-NMR spectrum.

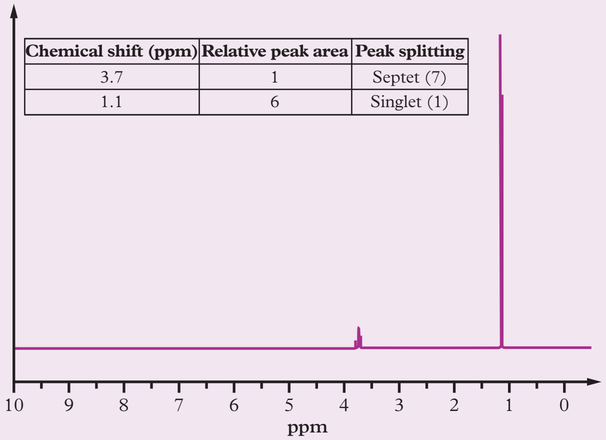
  d Describe what the ^{13}C-NMR spectrum would look like for this compound.

2 Describe the MS, IR, ^{13}C-NMR and ^{1}H-NMR spectra of the compound 2-methylpentan-3,3-diol, including peak area, fine splitting and shift, where applicable.



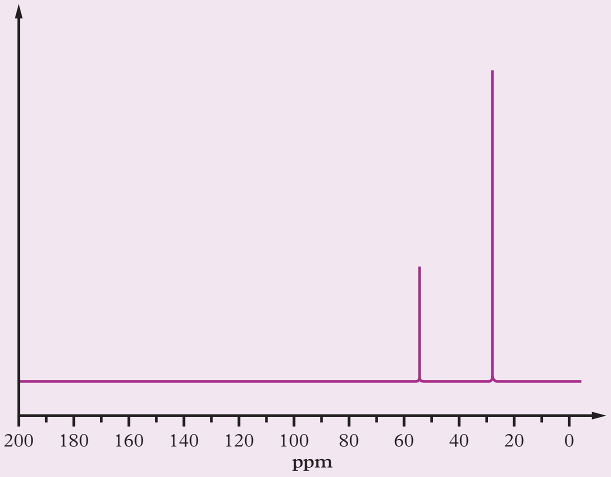
3 A compound contains carbon, hydrogen and chlorine and has a molar mass of 78.5 g mol^{-1}. The ^{1}H-NMR and ^{13}C-NMR spectra for this compound are shown below.

^{1}H-NMR spectrum:



Chemical shift (ppm)	Relative peak area	Peak splitting
3.7	1	Septet (7)
1.1	6	Singlet (1)

^{13}C-NMR spectrum:

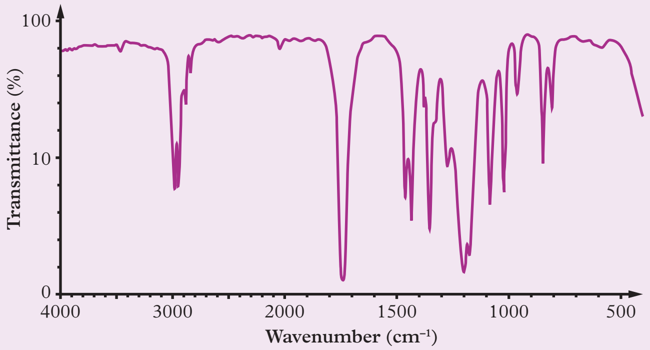


a If a mass spectrum were produced by this molecule, predict how many molecular ion peaks you would have, and what their *m*/*z* would be.

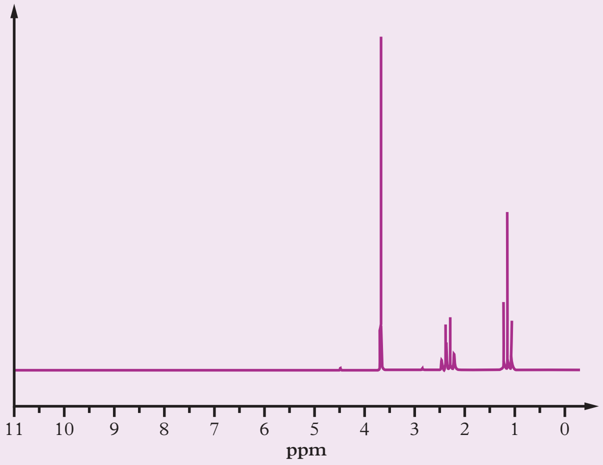
b Determine the name and structure of the organic compound.

4 The IR, ^{1}H-NMR and ^{13}C-NMR spectra for an unknown compound are shown below. The compound contains carbon, hydrogen and oxygen and has a molecular ion peak at $m/z = 88$ on a mass spectrum.

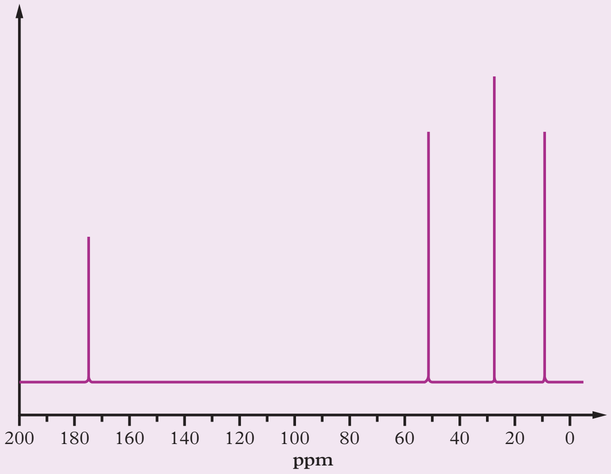
IR spectrum:



^{1}H-NMR spectrum:



^{13}C-NMR spectrum:



a Identify the characteristic functional group of the organic compound from the IR. Justify your answer.

b Identify the number of hydrogen environments in the unknown organic compound.

c Identify what the hydrogen environment at 1 ppm represents.

d Explain what the peak splitting in the ^{1}H-NMR might suggest.

e Identify the number of carbon environments in the unknown organic compound.

f Identify what the carbon environment at 174 ppm represents.

g Determine the semi-structural formula of the unknown organic compound.

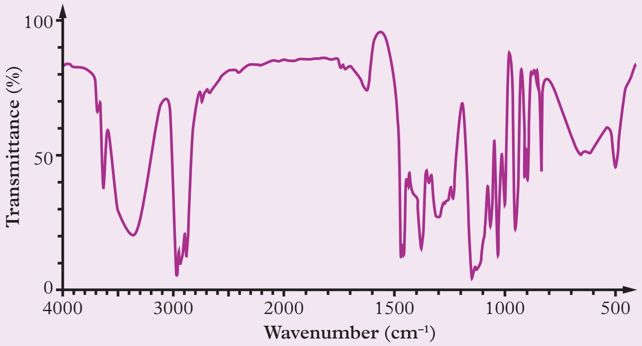
h Apply IUPAC naming rules to the compound you have determined.

5 The semi-structural formulas of four isomers with the molecular formula $C_5H_{12}O$ are shown below:

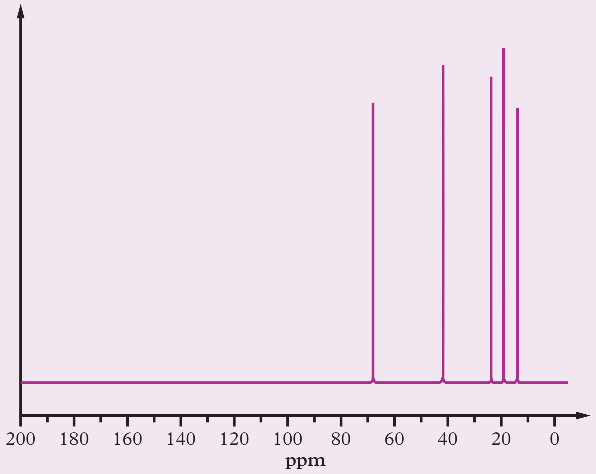
$CH_3CH_2CHOHCH_2CH_3$
$CH_3CH_2OCH_2CH_2CH_3$
$CH_3CHOHCH_2CH_2CH_3$
$(CH_3)_3COH$

One of the four compounds is run through IR and ^{13}C-NMR. The spectra produced are:

IR spectrum:



^{13}C-NMR spectrum:



a Explain why mass spectrometry would not be useful in identifying the compound.

b Determine which of the four compounds was analysed. Justify your answer.

Design and discuss

6 Discuss why one type of spectroscopy alone cannot always help chemists determine the structure of an organic compound, and why you would need to use a combination of different spectroscopy techniques.

7 Discuss which of the four instrumental analysis techniques (MS, IR, ^{1}H-NMR or ^{13}C-NMR) would give you the most information on its own. Justify your answer.

11.6

High-performance liquid chromatography

KEY IDEAS

In this topic, you will learn that:

- the components of a mixture can be separated using chromatography
- HPLC can be used for qualitative or quantitative analysis.

Chromatography is a common analytical technique used for both quantitative and qualitative analysis of compounds and mixtures. There are many different types of chromatography, including gas, paper and ion exchange chromatography. In this topic, we will look specifically at **high performance liquid chromatography** (HPLC), a highly **sensitive** analytical technique.

Principles of HPLC

Chromatography uses a mobile and a stationary phase to separate out individual components in a mixture. In HPLC, the sample mixture is dissolved into a suitable liquid **mobile phase** and injected into a column containing the **stationary phase**. Because different sample components will form different intermolecular interactions with the mobile and stationary phases, they will separate and leave the column at different times.

The amount and time taken (**retention time**) for each component in the sample to leave the column (or **elute**) is detected and expressed in a chromatogram for chemists to analyse. The liquid leaving the column (the **eluate**) can be collected and used, or disposed of as waste. This is summarised in Figure 1.

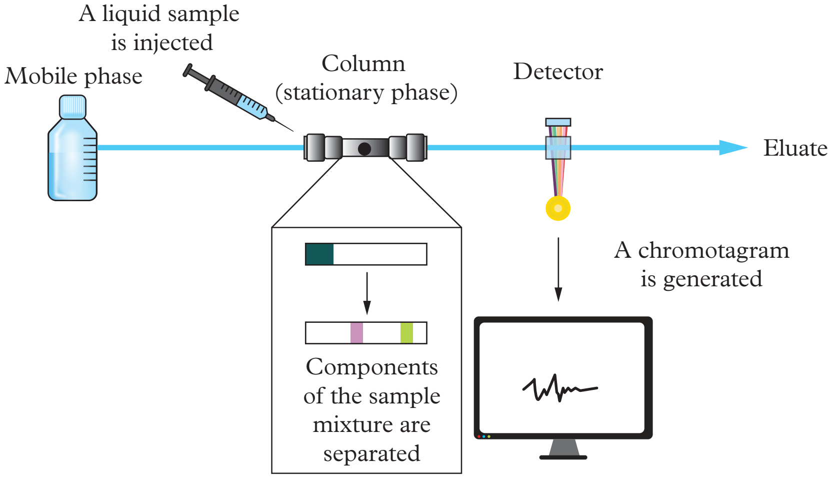


FIGURE 1 Overview of HPLC

high performance liquid chromatography (HPLC)
the technique used to separate, identify and calculate the amount of each component in a compound

sensitive
(in analytical chemistry) able to detect very low quantities of compound

mobile phase
the liquid that a sample is dissolved into, which then moves through the HPLC

stationary phase
the solid within the column of the HPLC onto which the components absorb

retention time (R_t)
the time that a component spends inside a chromatography column before it elutes

elute
exit of the chromatography column

eluate
the liquid that has left the column; may contain one or more sample components dissolved in the mobile phase

Study tip

You do not need to know the details about HPLC instrumentation and operation.

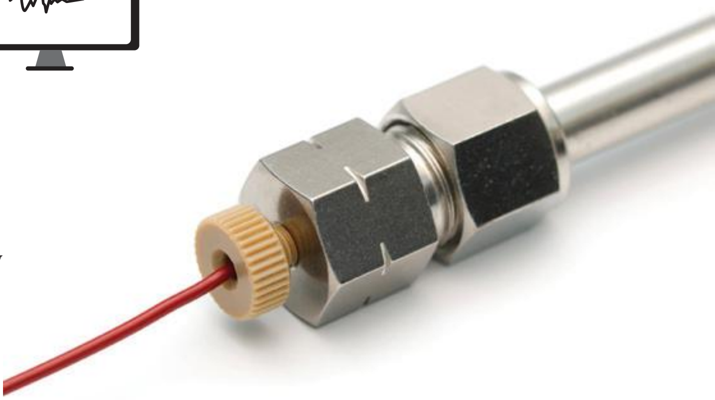


FIGURE 2 A chromatography column is tightly fitted to withstand high pressures.

Factors affecting separation

Remember from Unit 1 that the retention and separation of components in a sample mixture are based on two principles:

- **adsorption** to the stationary phase: the extent to which the components in the sample are attracted (by intermolecular forces) to the stationary phase
- **desorption** into the mobile phase: the extent to which the components in the sample dissolve into the mobile phase.

These processes are determined by the polarity of the components and the polarity of the phases, or the **like dissolves like** rule. Polar components are more attracted to polar stationary and mobile phases, and non-polar components are more attracted to non-polar stationary and mobile phases.

Figure 3 shows that the polar components (purple) elute more quickly from the column when a polar mobile phase is used. They have a shorter retention time because of their greater ability to dissolve better in the mobile phase. When a polar stationary phase is used, they will be retained longer because of their greater affinity for the stationary phase. Can you use this logic to explain what is happening to the non-polar (green) components?

adsorption
the process in which a component becomes attracted to the stationary phase

desorption
the process in which a component is released from the stationary phase and back into the mobile phase

like dissolves like
a rule stating that polar molecules will dissolve better in polar solvents and non-polar molecules will dissolve better in non-polar solvents

Study tip

Refer back to Chapter 6 in *Chemistry for VCE Units 1 and 2* if you need to revise the *like dissolves like* rule and affinity of compounds for the mobile and stationary phases.

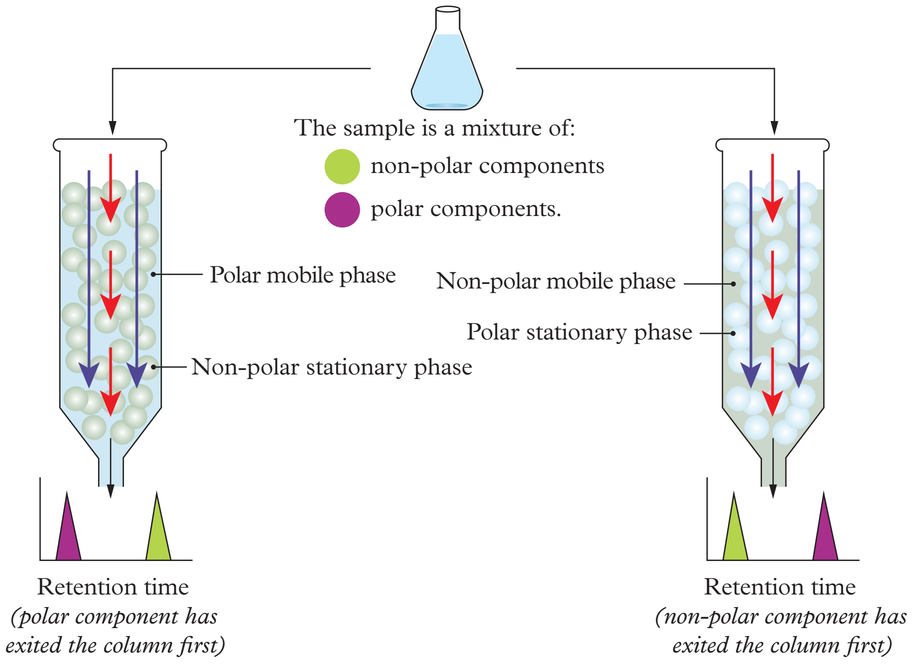


FIGURE 3 The effect of mobile and stationary phase polarity on retention times of polar and non-polar components

All of the HPLC conditions to consider, which will affect retention time, are:

- stationary phase: composition, polarity, surface area
- mobile phase: composition, polarity, flow rate, pressure
- length of the column
- temperature of the column.

Any changes to these conditions will affect the retention time. For example, increasing the temperature of the column will dissolve more of the sample into the mobile phase. You do not need to know how to explain the effect of changing all of these on retention time. But, you must know that retention times can only be compared against known retention times to identify a compound, if the HPLC conditions are the same.

FIGURE 4 HPLC samples are loaded into small vials.

Analysing HPLC chromatograms

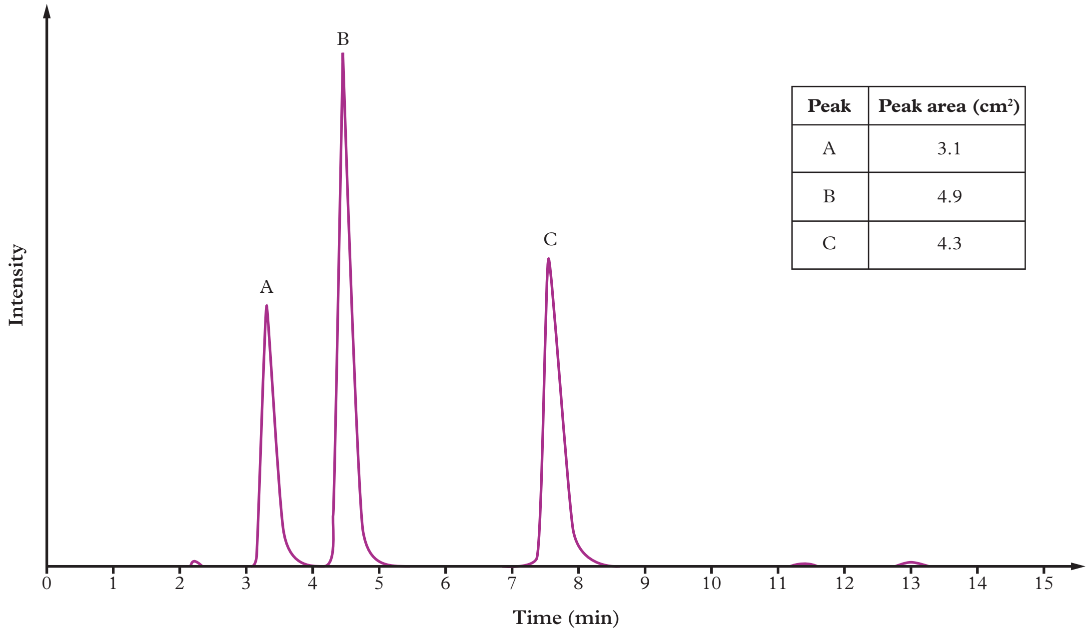
A HPLC chromatogram shows retention time (R_t) on the x-axis and signal intensity on the y-axis. They can be used for:

- qualitative analysis – to identify the components present in a sample
- quantitative analysis – to quantify the amount of a component present in a sample.

Qualitative analysis of chromatograms

You can identify the components in a sample by matching experimental retention times with known retention times found in the literature. Remember that this is only possible if the HPLC conditions are the same.

Figure 5 shows a chromatogram of a mixture consisting of three different sugar components. Each sugar has created a peak, labelled A, B and C. The retention times of each peak can be compared with known retention times (Table 1) to identify which peak belongs to which component.



Peak	Peak area (cm^2)
A	3.1
B	4.9
C	4.3

FIGURE 5 HPLC chromatogram

Study tip

Unusual or unexpected peaks in a chromatogram can also be used to identify the presence of contaminants.

Study tip

When reading retention times from a chromatogram, use a ruler to be exact, because some retention times can be close together.

To identify R_t on a chromatogram, you must read the retention time to find the middle point of the specific peak and then look to where that point sits along the x-axis.

- Peak A has a R_t of ~3.3 minutes, so it matches the retention time of fructose.
- Peak B has a R_t of ~4.5 minutes, so it matches the retention time of glucose.
- Peak C has a R_t of ~7.6 minutes, so it matches the retention time of sucrose.

The analysed sample contained fructose, glucose and sucrose, but not maltose or galactose.

TABLE 1 Known retention times of various sugars

Compound	R_t (min)
Maltose	10.08
Sucrose	7.55
Galactose	6.52
Glucose	4.45
Fructose	3.30

Quantitative analysis of chromatograms

You may have noticed that peak area was shown on the chromatogram in Figure 5. This value, expressed in cm^2, is proportional to the amount of substance present in a sample and is therefore important for quantitative analysis.

The quantity of a compound can be determined by running a set of known concentrations (standard solutions) of the compound and producing a **calibration curve**. This plots known concentrations against peak area. After this, we can find the peak area of the unknown samples on the calibration curve and match them to a concentration. But remember: the HPLC analyses must be run under the same conditions for you to compare values.

When selecting concentrations for your calibration curve, you need to have a rough idea of the concentration of your unknown compound. If you are significantly outside of the calibration curve concentration range, you would not be able to accurately determine the concentration. Make sure that your calibration curve has at least one data point above and one below this concentration.

Worked example 11.6 walks you through how to use a calibration curve to determine the concentration of sucrose in a sample.

calibration curve
a graph of known concentrations and peak areas from HPLC; used to identify an unknown concentration

11.6 Worked example
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11.6 Worked example
Video demonstration

11.6 SKILL DRILL

Evaluating the use of HPLC analysis for determining the concentration of fructose in soft drinks

Key science skill: Analyse and evaluate data and investigation methods

Two students decided to test the concentration of fructose in various soft drinks using HPLC. They developed a method to first test four known concentrations of high-purity fructose, so they could construct a calibration curve. They used the method below to test three cans of soft drink: cola, lemonade, and ginger beer.

Method

1 Transfer a 10.00 mL aliquot of the soft drink from the can into a 1.000 L volumetric flask.

2 Fill the volumetric flask up to the 1.000 L mark with deionised water and shake the flask.

3 Inject a sample of the diluted soft drink into the HPLC and run the analysis, using the same operating conditions used during calibration.

4 Repeat steps 1–3 for each of the other soft drinks.

Results

Fructose ($g\,L^{-1}$)	Peak area (cm^2)	Soft drink	Peak area (cm^2)
30	1355	Coca-Cola	5270
40	2610	Ginger beer	3122
50	3714	Lemonade	4191
60	4988		

Practise your skills

1 Create a calibration curve using the data the students collected.

2 Plot the three unknown samples on the calibration curve and determine the concentrations of fructose in each.

3 Discuss the error that the students made in selecting their known concentrations for the calibration curve.

4 Evaluate the precision and accuracy of their data.

5 Suggest changes to the method that could increase accuracy and precision.

Need help analysing and evaluating data and investigation methods? See Topic 1.8 (page 24).

11.6 CHECK YOUR LEARNING

Describe and explain

1 Explain how HPLC is used for quantitative analysis.

2 Describe the reason why HPLC must be performed under the same conditions to qualitatively analyse a sample.

3 Explain how two different substances could have the same retention time using the same conditions in the same HPLC.

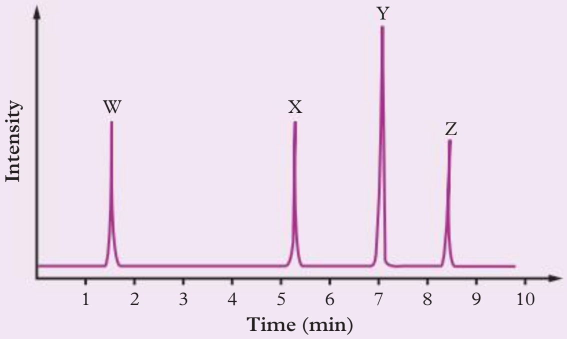
Apply, analyse and compare

4 A mixture containing ethanol, methane, butanoic acid and pentane is run through a HPLC system containing a polar stationary phase and a non-polar mobile phase.

a Predict the order in which they would exit the column. Justify your answer.

b In a separate experiment, the polarity of the mobile and stationary phases is swapped. Predict the results and compare this with the results in part **a**. Explain your reasoning.

5 Analyse the HPLC chromatogram shown.



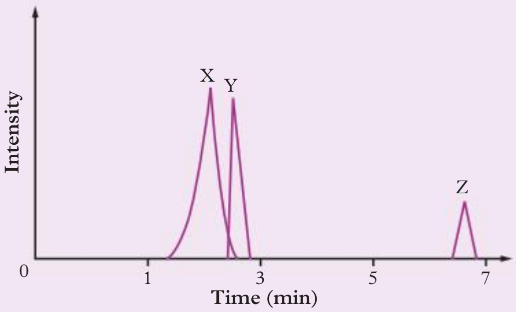
a If a polar stationary phase was used, describe the polarity of each of the components.

b Under the same conditions, hexane has a retention time of 7.1 minutes. Explain which peak would be hexane.

c On a different day, the same HPLC method was used to analyse a sample containing ethanol. The sample was found to have a retention time of 1.6 minutes. Explain whether we can conclude that peak W is ethanol.

Design and discuss

6 The following HPLC chromatogram is produced for a common energy drink. The retention times for possible components of the drink are shown in the table below.



Compound	R_t (min)
Caffeine	11.22
Fructose	3.84
Glucose	3.22
Sucrose	3.90
Taurine	7.62

a Suggest what compounds the peaks X, Y and Z could represent.

b Explain why it was hard to identify peaks X and Y.

c Suggest a change to the operating conditions of the HPLC that could help identify peaks X and Y.

7 The aspartame in a diet soft drink was quantified. Four samples of known concentrations of aspartame were analysed, followed by the unknown sample under the same conditions. The peak areas were recorded.

Concentration (mg/mL)	Peak area (cm^2)
0.20	0.74
0.40	1.51
0.60	2.23
0.80	2.85
Diet soft drink sample	1.90

a Create a calibration curve using the data provided in the table.

b Determine the amount of aspartame from the calibration curve.

c Calculate the amount of aspartame in the 375 mL can of energy drink.

11.7 Applications of laboratory and instrumental analysis

KEY IDEAS

In this topic, you will learn that:

- laboratory and instrumental analytical techniques are used in various industries to identify both desired and undesired compounds.

In this topic, we will look at the application of laboratory and instrumental methods in society to measure product purity and confirm the presence of organic compounds in a product.

Determining product purity

percentage purity the percentage of the desired product contained in a product

Product purity is very important in many of the consumer products we buy and use daily. For example, if medicines contained harmful impurities, this may counteract some of the active ingredients or cause harmful side effects. Chemists and manufacturers must determine the percentage of the desired chemical in a sample, or **percentage purity**, of their products. This is quite simple to test and calculate. The following equation is used:

$$\%\ \text{purity} = \frac{\text{mass of desired product}}{\text{total mass of sample}} \times 100$$

11.7A Worked example
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11.7A Worked example
Video demonstration

To measure percentage purity, you must first quantitatively analyse the sample to find out how much of the desired product it contains. This can be done using one of the many quantitative analytical techniques we have looked at in Chapters 10 and 11, like volumetric analysis, HPLC or distillation. Once the mass of the desired product is determined, the equation above can be used. Worked example 11.7A shows how to calculate the percentage purity of a product.

The mass of desired products is determined by various chemical analysis techniques. Let's explore the chemical analysis of three different example consumer products: cosmetics, food products, and wine.

FIGURE 1 Cosmetics and toiletries such as sunscreen must undergo rigorous chemical tests before they can be put on supermarket shelves.

Chemical analysis of cosmetics

Every time you wash your hands, or put on makeup or sunscreen, you are applying a variety of chemicals onto your skin. Toiletries, cosmetics, and their ingredients are all highly regulated by several Australian government agencies. These products are all subjected to rigorous testing before they can be sold.

Chemical analysis of skin and beauty care products must be done to make sure that:

- finished products are safe for consumers
- if restricted chemicals are used, their quantities are within safe limits
- raw materials are of high quality
- products adhere to mandatory labelling requirements in Australia.

11.7
PRACTICAL: FIELDWORK
How is instrumental analysis used in industry? Go to page **523**.

FIGURE 2 Cuvettes used for UV-Vis spectrophotometry

These products can be analysed using many different chemical analysis methods, including:

- chromatography
- calorimetry
- distillation
- volumetric analysis
- UV-Vis spectrophotometry
- mass spectrometry (MS)
- infrared spectroscopy (IR).

In this topic, we will only look at qualitative analysing of cosmetics using IR spectroscopy and distillation, and quantitative analysis using volumetric analysis. Have a think about how the other techniques can be used, too.

Qualitatively analysing purity of cosmetics using IR

IR spectroscopy can be used to assess the quality and purity of raw materials used in skin and body care products. These raw materials are sourced from various Australian and overseas suppliers. Before using the raw material to create their product, chemists will analyse the material using IR and compare it with the spectrum of a pure reference sample. This is done to confirm that the identity and purity of the material matches what the original supplier has claimed. An example is shown in Figure 3.

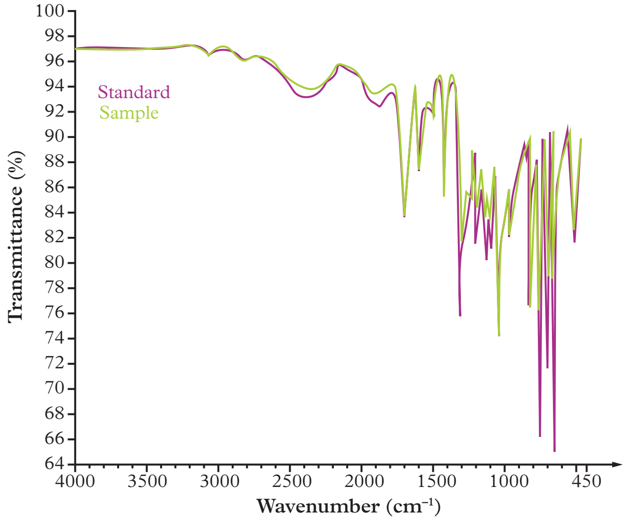

FIGURE 3 Overlaid IR spectrum (purple) of a known nicotinic acid reference standard with a nicotinic acid stock purchased from a supplier (green). This stock contained nicotinic acid.

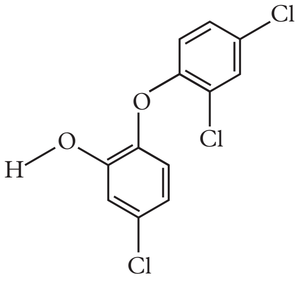

FIGURE 4 The structure of triclosan

IR is also used to analyse final products and make sure that they meet Australian regulatory standards. This can determine whether products contain any restricted chemicals and check if they are within or outside the legal limits.

For example, the chemical triclosan (Figure 4) is an antibacterial and antifungal agent added to deodorant, soap, and a variety of cosmetics. Studies have shown that triclosan is linked to a range of negative health and environmental effects. However, in Australia AICIS (the regulating body for industrial chemicals in Australia) produced a full risk assessment of triclosan and found ‘no cause for public concern’ if the maximum concentration in consumer products is 0.3%.

Any cosmetic that contains over 0.3% triclosan must have the word 'poison' on the label, which is not a great selling feature to win over customers. Analytical chemists now use IR spectroscopy to detect triclosan and other regulated chemicals (Figure 5).

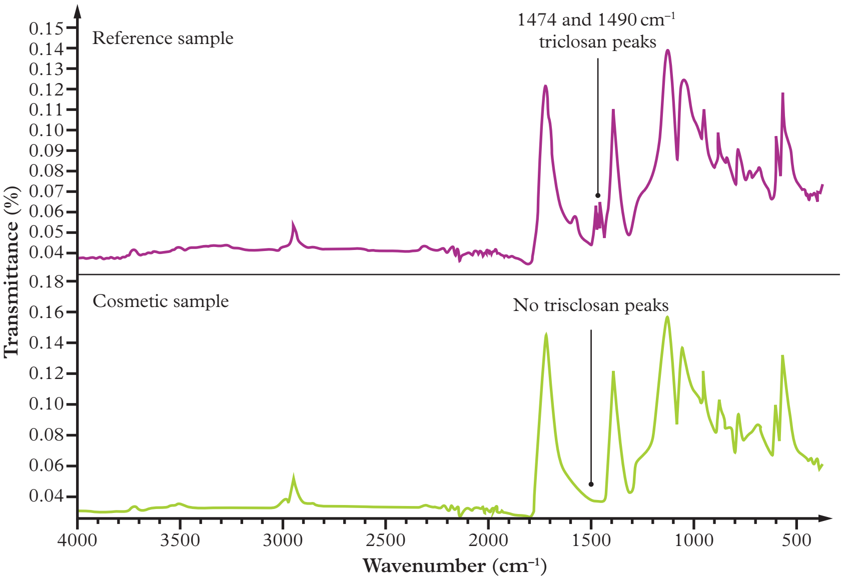

FIGURE 5 IR spectra of a reference sample containing triclosan and a cosmetic sample with unknown triclosan content

Though IR cannot reliably determine the quantity of triclosan or other restricted chemicals in a sample, it can quickly identify if the chemical is present. Often, chemists use software that compare analysed samples with reference spectra in under 30 seconds. Once a chemist has determined the presence of a restricted chemical (or any chemical they are looking for) using IR, they can then quantitatively analyse the sample using other chemical analysis methods.

Qualitatively analysing purity of cosmetics using distillation

Essential oils are widely used in cosmetics, perfumes and household products, but are mainly used for their pleasant aromas. Including essential oils in beauty products is always seen as attractive and a marketing advantage, but essential oils come with a higher price tag. Because of this, essential oils are often **counterfeited**. Sometimes synthetic oils are used to 'bulk-out' or replace essential oils, and are sold for the same price.

counterfeited
an inauthentic imitation of a substance

To verify the purity of an essential oil, many different analytical tests can be performed. They are often the same analytical tests used to confirm purity of raw materials in cosmetics. For example, IR spectra can be compared to find out if an essential oil is counterfeit.

Distillation can also be used to determine the boiling point of an essential oil. This can be compared against known values to confirm the identity of the oil. For example, an essential that boils at around 176°C is likely to be rosemary, whereas if it boils at a much higher temperature, like 204°C, it may be lavender.

FIGURE 6 Lavender oil

Quantitatively analysing compounds in cosmetics using volumetric analysis

FIGURE 7 Potassium permanganate reacting with hydrogen peroxide

In recent years, the hand sanitiser industry has expanded enormously. In 2020, the spread of COVID-19 created a hand sanitiser shortage, which allowed some subpar sanitisers to reach the market.

According to the World Health Organisation (WHO), hand sanitisers must contain a combination of ethanol or isopropyl alcohol, hydrogen peroxide and glycerol to be effective.

Lots of testing was done in the early 2020 hand sanitiser boom to check that products on the shelves were up to the standards of the Therapeutic Goods Administration (TGA) of Australia. In fact, one company was fined $25 200 for breaching the *Therapeutic Goods Act 1989* with the hand sanitisers they were selling.

One commonly used test to check that the hydrogen peroxide, H_2O_2, concentration is within acceptable limits is a self-indicating redox titration of H_2O_2 with potassium permanganate, $KMnO_4$.

$KMnO_4$ causes the oxidation of H_2O_2 according to the following redox equation:

$$5H_2O_2(aq) + \underset{\text{purple}}{2MnO_4^-(aq)} + 6H^+(aq) \rightarrow 5O_2(g) + \underset{\text{colourless}}{2Mn^{2+}(ag)} + 8H_2O(l)$$

11.7B Worked example
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11.7B Worked example
Video demonstration

Therefore, $KMnO_4$ can be used to determine the amount of H_2O_2 contained in the hand sanitiser solution. The $KMnO_4$ starts as a dark purple colour and, as it is reduced, it becomes colourless. As soon as there is no hydrogen peroxide left in the solution, the solution will change back into purple.

Worked example 11.7B walks you through how to calculate the concentration of H_2O_2 in hand sanitiser, based on results from a redox titration.

Chemical analysis of food products

FIGURE 8 Food and drink products are routinely analysed to ensure quality.

Foods are very complex mixtures of chemicals. For nutritional and safety reasons, we want the contents of the food product to match what is stated on the label. Food Standards Australia New Zealand (FSANZ) develop the food standards, labelling requirements and food safety standards for Australia and New Zealand. FSANZ continually test different foods to make sure that they are up to these standards. HPLC, in combination with mass spectrometry, is one of the leading techniques to guarantee food authenticity in order to address food safety issues.

HPLC can be used qualitatively to check that food products contain the correct (or incorrect) substances, and also quantitatively, to make sure that labels have the accurate amounts of nutrients on them.

Qualitatively analysing food products using HPLC

In 2018, a large honey scandal occurred in Australia. It was found that 27% of commercial honey brands in Australia had 'watered down' their honey with corn syrup, sugar cane or other products. Australia, being the fourth largest honey exporter in the world, had its reputation on the line.

One test used to determine the purity of honey was HPLC analysis. The honey was then compared to a reference sample. Figure 9 shows chromatogram sections of an altered honey sample against an authentic honey sample.

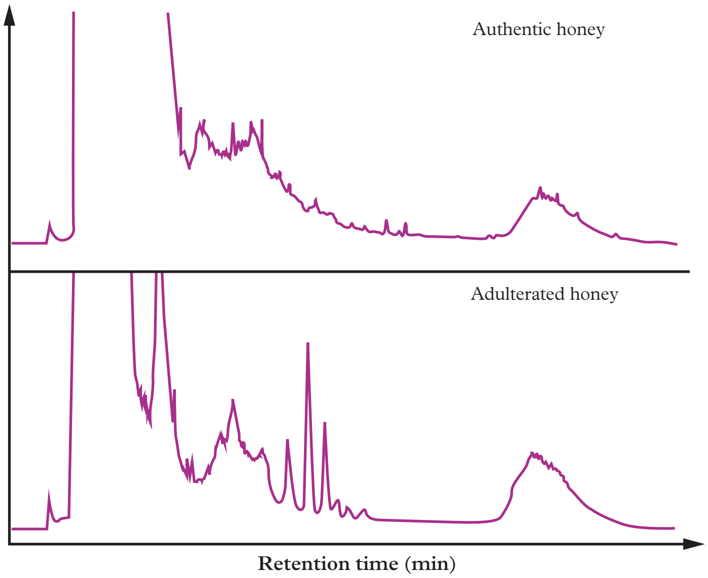


FIGURE 9 Chromatograms of authentic and adulterated honey

Study tip

Remember, to be able to directly compare chromatograms, the samples must be run through a HPLC using the same conditions.

Chemists can also analyse the sugars in honey samples using HPLC to check if all sugars present are naturally occurring sugars in honey. The chromatogram in Figure 10 determined that flower honey was adulterated with rice syrup, an oligosaccharide. Natural honey does not contain oligosaccharides and would therefore not include a peak at R_t = 1100 min.

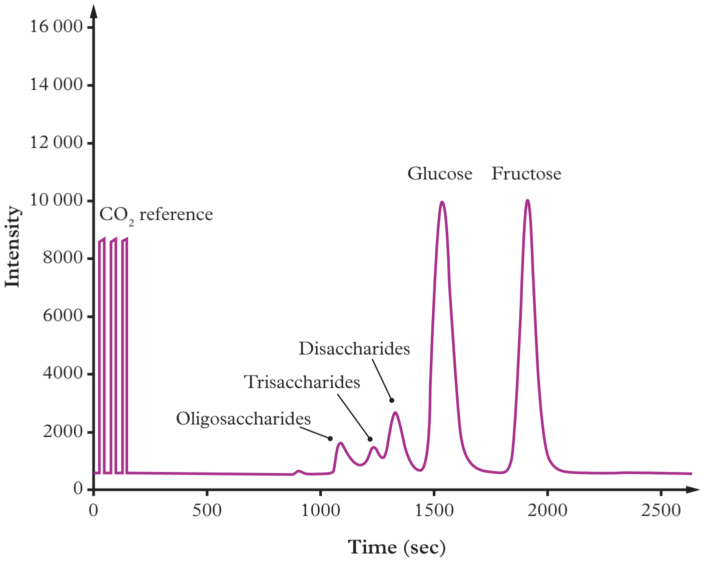


FIGURE 10 HPLC chromatogram of an adulterated honey sample

FIGURE 11 HPLC was used to test the purity of a honey sample.

Quantitatively analysing food products using HPLC

HPLC can also be used to quantify microplastics consumed by fish in the ocean. In the chromatogram for prawn meat (Figure 12), the highlighted peaks correspond to common plastics found in the prawn.

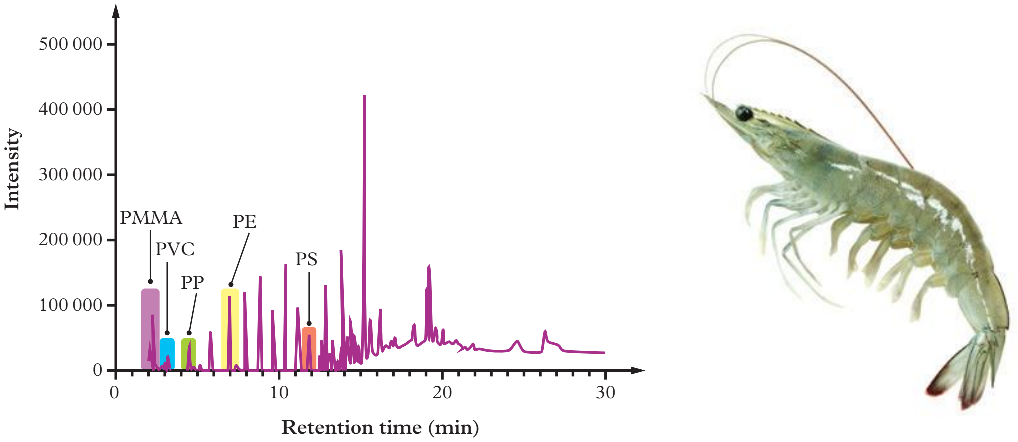


FIGURE 12 Prawn meat chromatogram with highlighted plastics: poly(methyl methacrylate) (PMMA), polyvinyl chloride (PVC), polypropylene (PP), polyethylene (PE) and polystyrene (PS)

In one analysis of different seafood varieties, it was found that different species accumulated different amounts of plastics. They concluded this by analysing known concentrations of the liquified plastic samples and producing a calibration curve. They then plotted the seafood sample peak areas and determined the unknown concentration from the calibration curve (Figure 13).

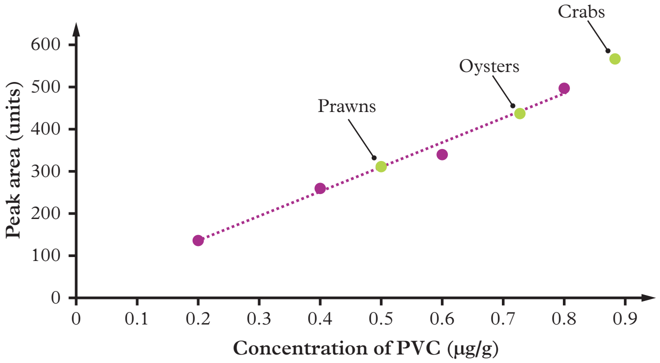


FIGURE 13 Calibration curve of known concentrations of PVC and the calculated concentrations in prawns, crabs and oysters

Study tip

All of the analytical techniques explored in Chapters 10 and 11 can be used in combination to qualitatively and quantitatively analyse organic compounds.

Ideally, there would be no plastic present in the seafood that we consume, but microplastics are present in many fish species. Scientists need to continue monitoring and analysing seafood to determine what levels are safe for humans to consume. Hopefully, with more conscious recycling and plastic use, we can reduce microplastic levels present in fish species and in the ocean. This is also a good opportunity to discuss the sustainable development goals.

FIGURE 14 Microplastics are found in high amounts in crabs.

Chemical analysis of wines

Wines and other alcohols must be analysed to ensure that they are safe for people to consume. It is also used to check that wines for sale in bottle shops and restaurants adhere to strict labelling and regulations around additives. The regulations for wines made and sold in Australia are very strict, so winemakers need to analyse many chemicals in their wines throughout the production process.

FIGURE 15 Before wine and other alcohols can be sold to the public, they must undergo series of chemical analyses.

Quantitatively analysing alcohol content using distillation

Laboratory and instrumental analysis techniques can be used to find the percentage of ethanol (alcohol content) in wines. During the wine-making process, alcohol content is measured in wines both during and after fermentation.

The simplest technique is to use fractional distillation to separate volatile and non-volatile components. The fraction containing ethanol can then be tested using a **hydrometer** to check for the alcohol content. Hydrometers are instruments that are highly calibrated and use density to give a measurement of alcohol in wine. They provide a quick and easy way to measure alcohol content, and most wineries will use this method if they do not have access to more costly and accurate ways of measuring alcohol content. A more precise measurement is often done prior to bottling wines, and that measurement is shown on the label.

hydrometer
an instrument used to measure the density of a substance

Quantitatively analysing alcohol content using IR

The more precise measurement for alcohol content is to use one of two different versions of IR spectroscopy for quantitative analysis of alcohol percentage in wine: near-infrared spectroscopy (NIR) or mid-infrared (MIR) spectroscopy. In both techniques, samples are inserted into either an NIR or MIR spectrophotometer and the amount of radiation absorbed is measured. Values obtained are checked against reference samples to determine concentration. Both techniques are fast, use a very small sample of wine, and are very precise methods of alcohol analysis.

FIGURE 16 Hygrometers are used to measure the density of substances.

Quantitatively analysing sulfur dioxide in wines using volumetric analysis

Sulfur dioxide, SO_2, is one of the common preservative additives in wines. It is used to prevent the growth of bacteria and undesirable wild yeasts that grapes can naturally have on their skin and to prevent wines from oxidising.

In Chapter 9, you looked at the oxidation of alcohols and how they become carboxylic acids. Figure 18 shows that there is an intermediate step where alcohols are partially oxidised to form aldehydes before being fully oxidised to form carboxylic acids.

FIGURE 17 Wine is made from the fermentation of grapes.

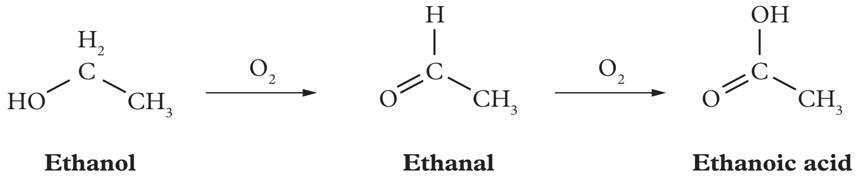


FIGURE 18 Ethanol will react with oxygen partially to form ethanal and then further to form ethanoic acid.

While ethanoic acid and ethanal are both very useful compounds, neither is desirable in wine. Ethanoic acid is essentially vinegar and ethanal can have a strong smell, similar to nail polish remover. Both molecules affect the overall taste of the wine. Sulfur dioxide is added to prevent oxidation from occurring and limit the presence of these molecules in wines. It reacts with oxygen molecules so they cannot oxidise ethanol.

aeration
the process in which a substance is exposed to air

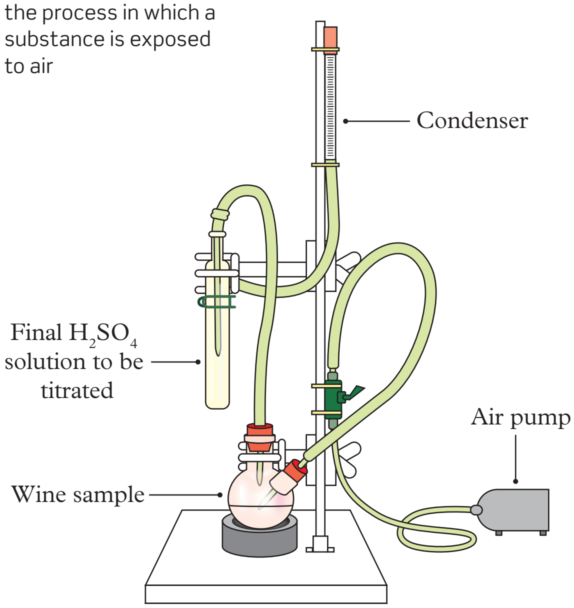


FIGURE 19 The SO_2 in wine can be analysed using aeration and oxidation. After this step, the final H_2SO_4 solution is analysed using titration.

Winemakers and wine laboratory technicians need to keep a constant eye on levels of SO_2. Too little SO_2 will cause oxidation to occur, but too much can strip wine of its colour and ruin its taste. Winemakers analyse the amount of SO_2 in their wines and calculate how much they must add or remove to stay within the optimum range. A simple laboratory analysis that is used to test for SO_2 dioxide in wines is the **aeration**, oxidation and acid–base titration method.

Wine samples are acidified and then aerated into a solution of hydrogen peroxide. This oxidises the sulfur dioxide to sulfuric acid, H_2SO_4. The sulfuric acid formed is then titrated with standardised sodium hydroxide, NaOH, according to the following equation:

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

The amount of NaOH required to reach end point is proportional to the amount of SO_2 in the wine, and the concentration of SO_2 in the wine can be calculated using stoichiometry.

11.7 CHECK YOUR LEARNING

Describe and explain

1 Explain why product purity is important in pharmaceuticals.

2 Explain why cosmetics must be tested using analytical and instrumental techniques before being sold.

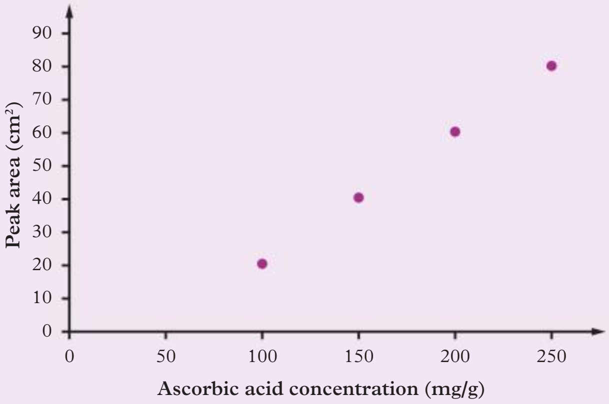
3 The equation for the redox titration between H_2O_2 and $KMnO_4$ to determine hydrogen peroxide concentration in hand sanitiser is:

$$5H_2O_2(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5O_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$$

Construct the two redox half-equations and explain how a redox titration can occur.

Apply, analyse and compare

4 The packaging for a vitamin C tablet says that each 2.30 g tablet contains 400 mg of ascorbic acid. The tablet was analysed using HPLC and the peak produced on the chromatogram had a peak area of 43 cm^2. A calibration curve of known concentrations of ascorbic acid was then prepared as shown.



a Analyse the calibration curve and plot the vitamin C tablet on it to identify the ascorbic acid concentration.

b Determine the amount of vitamin C in the tablet that was analysed.

c Calculate the percentage purity of the ascorbic acid in the vitamin C tablet.

5 A redox titration with potassium permanganate was performed to calculate the concentration of H_2O_2 in a hand sanitiser. A 10 mL sample was diluted in 250 mL and 25 mL aliquots were titrated against 0.010 M $KMnO_4$. Four titres of 14.13, 14.28, 14.21 and 14.15 mL were obtained. Calculate the %(*m*/*v*) of hydrogen peroxide in the hand sanitiser.

Design and discuss

6 Discuss why it is important for fuels used in transport vehicles to have a high percentage purity.

CHAPTER 11 Review

Chapter summary

11.1
- Mass spectrometry is used to identify the mass-to-charge ratio (m/z) of molecular ions.
- The molecular ion peak on a mass spectrum is equal to the molar mass of the compound. The base peak is the ion with the highest relative intensity.

11.2
- Infrared spectroscopy is used to identify functional groups within an organic compound.
- All functional groups produce unique peaks on a spectrum as different bonds will rotate and vibrate differently, and therefore absorb different amounts of IR radiation.

11.3
- NMR uses electromagnetic radiation to flip the spin states of nuclei, and measures the different amounts of energy required to do this.
- The number of peaks in a ^{13}C-NMR spectrum is equal to the number of carbon environments in the compound.
- ^{13}C chemical shift is given in ppm and can be referenced in the data book or from Table 1 in Topic 11.3.

11.4
- The number of peaks in a ^{1}H-NMR spectrum is equal to the number of hydrogen environments in the compound.
- Fine splitting is caused by neighbouring hydrogen environments and uses the $n + 1$ rule to determine the number of fine splitting peaks.
- ^{1}H chemical shift is given in ppm and can be referenced in the data book or from Table 1 in Topic 11.4.

11.5
- MS, IR, ^{13}C-NMR and ^{1}H-NMR each give specific information about the bonds and structures of organic molecules.
- Combining the different spectroscopy techniques can reveal the identity of an unknown compound.

11.6
- HPLC can be used for qualitative analysis by determining experimental retention times and comparing them with known retention times.
- HPLC can be used for quantitative analysis by using a calibration curve to find the concentrations of unknown substances.

11.7
- Laboratory and analytical techniques are used in a variety of industries to identify desired or undesired products and percentage purity.

Key formulas

Amount of splitting on the neighbouring peak (for ^{1}H-NMR data)	$n + 1$
Percentage purity	$\%\text{purity} = \frac{\text{mass of desired product}}{\text{total mass of sample}} \times 100$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as ‘Confidently’, ‘Mostly’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Mostly	Not really	Revision link
• describe the applications of MS	☐	☐	☐	Go back to Topic 11.1
• interpret qualitative and quantitative MS data, including the identification of molecular ion peak, determination of molar mass and identification of simple fragments	☐	☐	☐	Go back to Topic 11.1
• analyse IR data using characteristic absorption bands to identify bond types	☐	☐	☐	Go back to Topic 11.2
• analyse low resolution ^{13}C-NMR spectra using chemical shift values to determine the number and nature of different carbon environments	☐	☐	☐	Go back to Topic 11.3
• analyse low and high resolution ^{1}H-NMR using chemical shift values, integration curves and peak splitting patterns to determine the structure of organic compounds	☐	☐	☐	Go back to Topic 11.4
• use the $n + 1$ rule to deduce the number and nature of different proton environments	☐	☐	☐	Go back to Topic 11.4
• use a combination of MS, IR, ^{1}H-NMR and ^{13}C-NMR data to deduce the structures of simple organic compounds	☐	☐	☐	Go back to Topic 11.5
• analyse HPLC chromatograms using calibration curves and retention times to determine the concentration of organic compounds in solution	☐	☐	☐	Go back to Topic 11.6
• describe the roles and applications of laboratory and instrumental analysis in: the identification of organic compounds or functional groups in isolation or within a mixture, and product purity	☐	☐	☐	Go back to Topic 11.7

Revision questions

Multiple choice

1 A ^{1}H-NMR spectrum of 2-chlorobut-1-ene produces a ^{1}H-NMR spectrum with three peaks. Which best describes the splitting pattern?

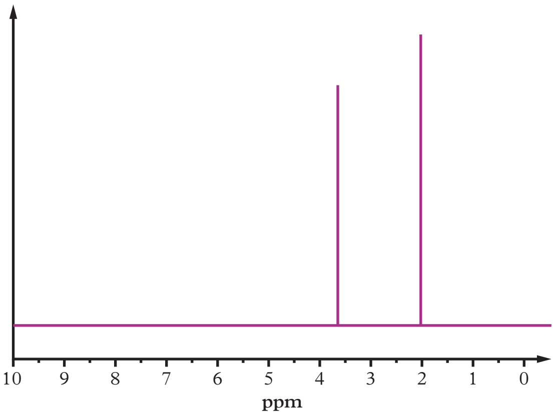
A 3 singlets
B 2 doublets and a triplet
C A singlet, a doublet and a quartet
D A singlet, a triplet and a quartet

2 Which of the following statements is incorrect?

I ^{1}H-NMR and ^{13}C-NMR use electromagnetic radiation to determine the number of hydrogen and carbon environments respectively.
II IR is used to find the molar mass of a compound.
III ^{1}H-NMR can be used to determine the functional group of a compound.
IV All three techniques measure the wavenumber in ppm absorbed by the molecule.

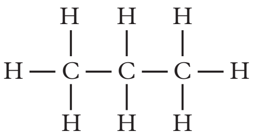
A II only
B III only
C I and II
D II and IV

3 A ^{1}H-NMR spectrum is shown.



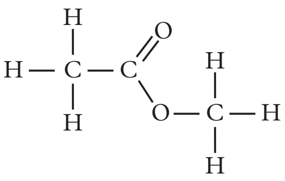
Which compound does it belong to?

A

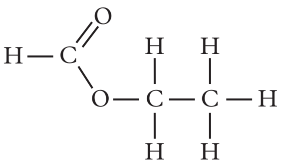


$$H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$$

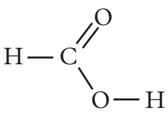
B



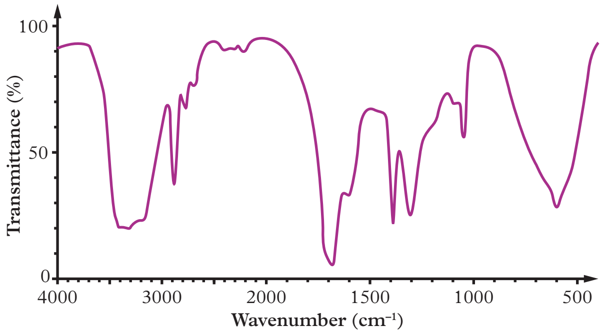
C



D



4 An IR spectrum is shown below.



Which compound does it belong to?

A Methanoic acid

B Methane

C Chloromethane

D Methanamide

5 An unknown organic molecule is analysed using ^{1}H-NMR and ^{13}C-NMR. The data is summarised in the table below.

^{1}H-NMR	^{13}C-NMR
3 peaks: a singlet, a doublet and a septet	2 peaks

What is the compound?

A Propan-1-ol

B Propan-2-ol

C 2-methylpropanoic acid

D Methyl methanoate

6 A compound has the molecular formula $C_3H_6O_2$. Which of the following would be the most useful in identifying the molecule?

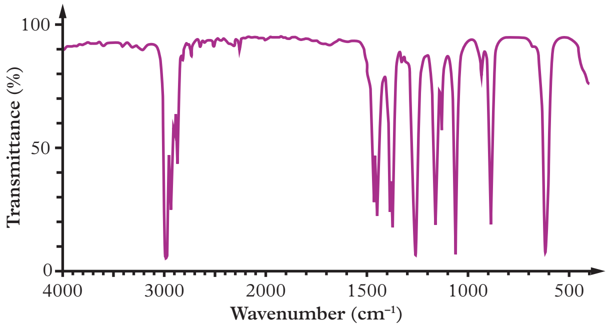
A Volumetric analysis

B Nuclear magnetic radiation

C Infrared spectroscopy

D Gravimetric analysis

7 An IR spectrum is shown below.



Which compound does it belong to?

A 2-chloropropane

B Propan-2-ol

C Propanoic acid

D Propan-2-amine

8 Which of the following compounds would *not* produce a triplet and a quartet in its ^{1}H-NMR spectrum?

A Ethanol

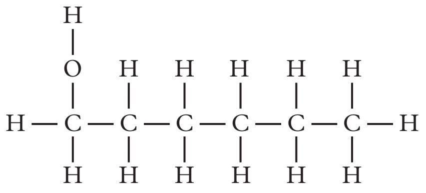
B Ethyl propanoate

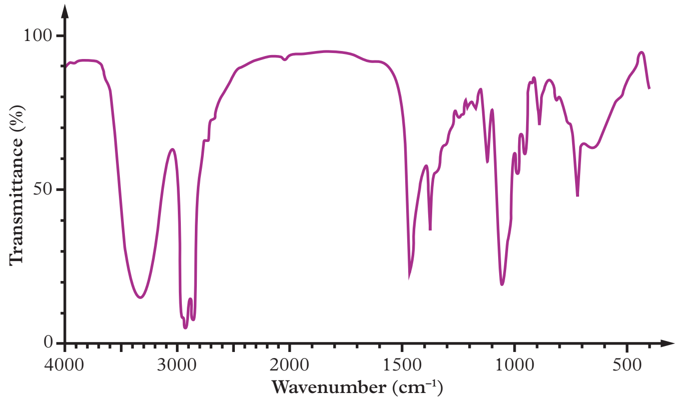
C Ethene

D Chloroethane

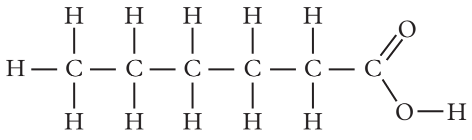
9 Which of the following is not the correct IR spectrum for the molecule shown?

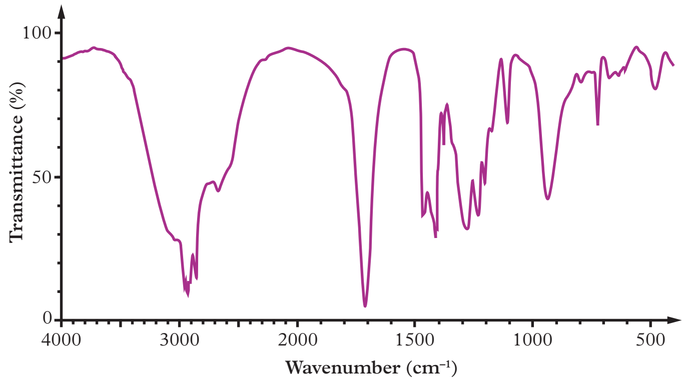
A



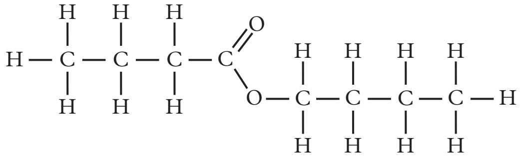


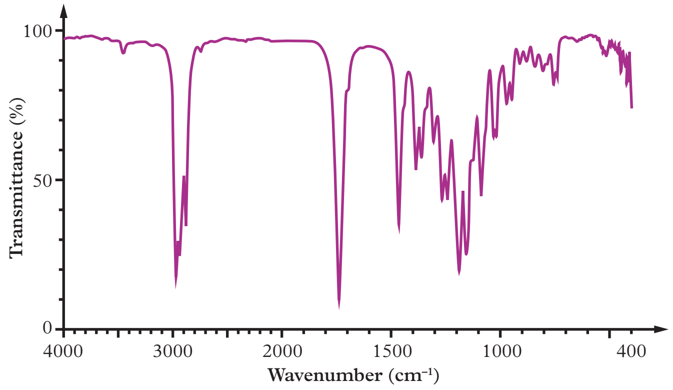
B



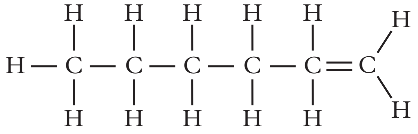


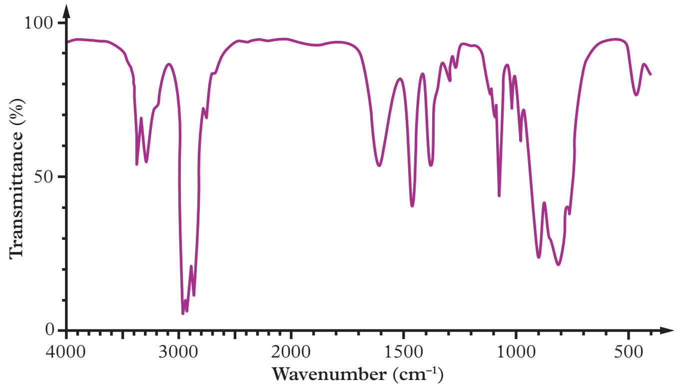
C





D





10 Which of the following IR, ^{1}H-NMR and ^{13}C-NMR results would be correct for propanol?

	IR description	Number of carbon environments	Number of hydrogen environments
A	Smooth broad strong peak at 2500–3500 cm^{-1}	3	3
B	Smooth broad strong peak at 3200–3600 cm^{-1}	3	4
C	Smooth broad strong peak at 2500–3500 cm^{-1}	2	3
D	Smooth broad strong peak at 3200–3600 cm^{-1}	2	4

Short answer

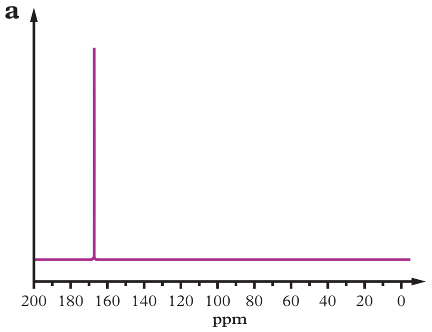
Describe and explain

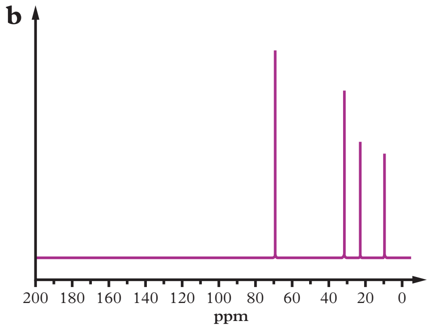
11 Describe the absorption bands that would be present on an IR spectrum for the following functional groups.

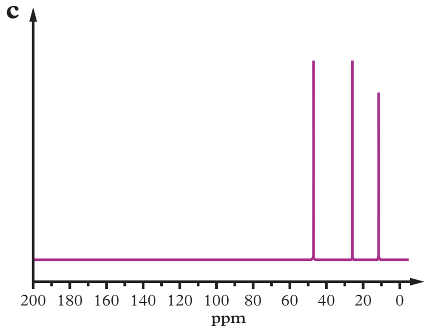
- **a** Amine
- **b** Carboxylic acid
- **c** Alcohol
- **d** Ester

12 Explain why TMS is used in NMR.

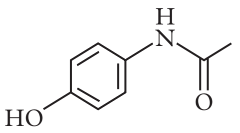
13 Describe each chemical shift in the following spectra and suggest the carbon environments that each matches.



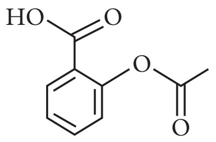




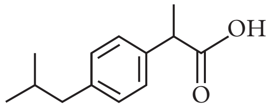
14 A mixture (containing the three drugs shown below) is analysed using a HPLC with a polar mobile phase and non-polar stationary phase. Predict the order in which the drugs will appear on the chromatogram. Explain your reasoning.



Paracetamol

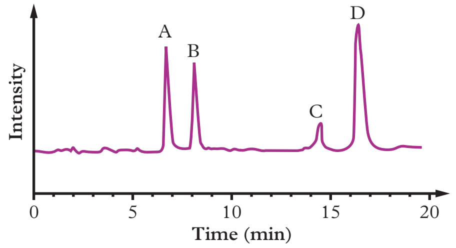


Aspirin



Ibuprofen

15 The following chromatogram was produced from running a dissolved breakfast cereal sample through a HPLC with a non-polar stationary phase and a polar mobile phase.



a Identify the retention time of compound D.

b List compounds A–D in order of increasing polarity. Justify your reasoning.

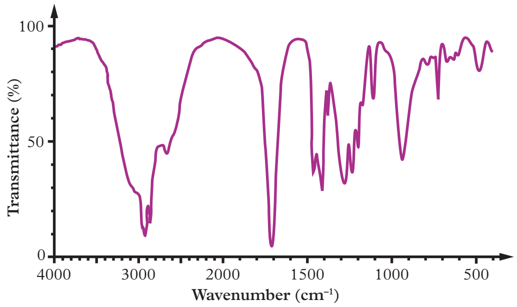
c Explain what information the peak area of A–D gives you from this chromatogram.

16 Explain how IR can be used to analyse the quality of cosmetics.

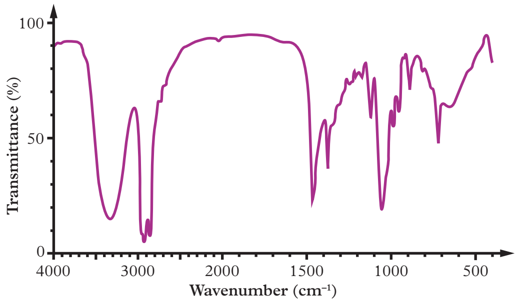
17 Describe the peaks you would expect to see in the ^{13}C-NMR and ^{1}H-NMR spectrum for cyclohexane.

18 Describe the absorption bands present in the functional group region and their wavelengths in the following IR spectra:

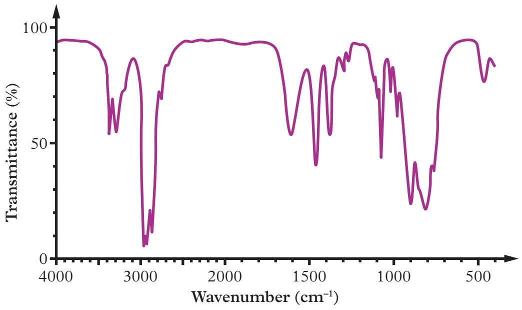
a



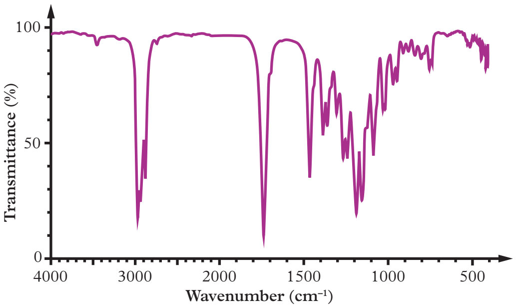
b



c



d



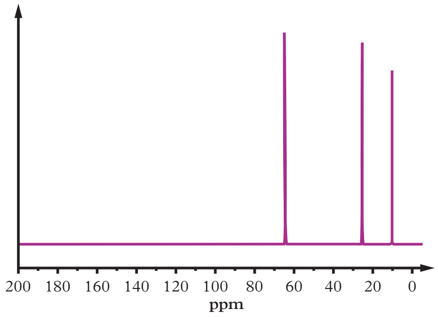
19 Describe the ^{1}H-NMR spectrum you would expect to see for propane. Include chemical shift, peak area and peak splitting.

Apply, analyse and compare

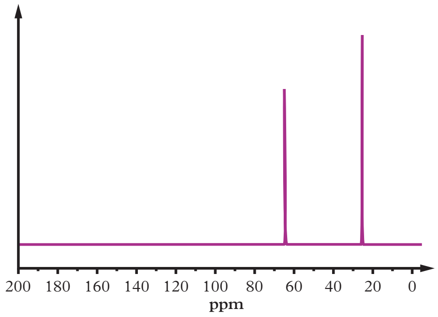
20 A redox titration with potassium permanganate was performed to calculate the concentration of H_2O_2 in a hand sanitiser. A 20 mL sample of hand sanitiser was diluted in 100 mL and 10 mL aliquots were titrated against 0.030 M $KMnO_4$. Four titres of 15.18, 15.23, 15.15 and 15.21 mL were obtained. Calculate the %(m/v) of hydrogen peroxide in the hand sanitiser.

21 Analyse the following ^{13}C-NMR spectra **a**–**d** and match them with the four compounds propanoic acid, methyl ethanoate, propan-1-ol and propan-2-ol.

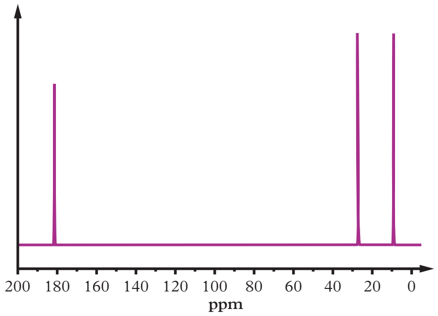
a



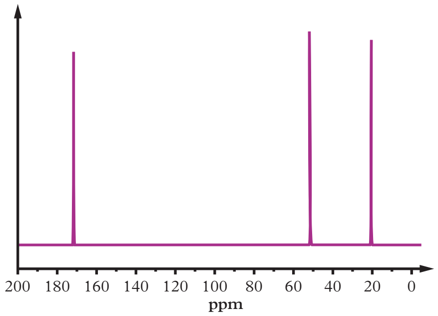
b



c

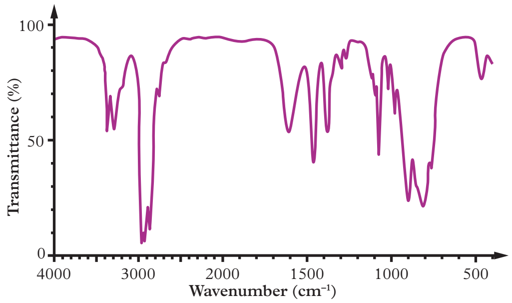


d

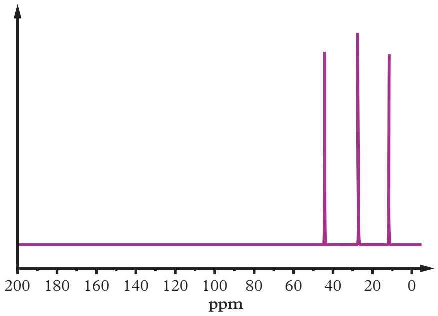


22 Analyse the IR and ^{13}C-NMR spectra below and suggest an organic compound that would produce them both.

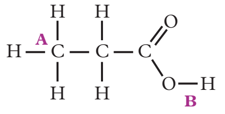
IR spectrum

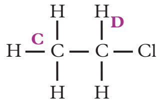


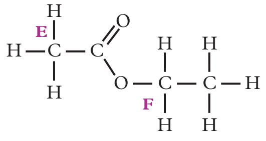
^{13}C-NMR spectrum

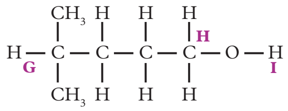


23 The following four compounds were analysed using ^{1}H-NMR. The H environments are labelled A–J.









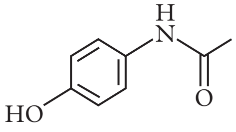
Match the G environment in the four compounds with the descriptions below.

- **a** A quartet at δ = 4.8 ppm
- **b** A singlet at δ = 3.0 ppm
- **c** A triplet at δ =4.5 ppm
- **d** A singlet at δ =12.0 ppm
- **e** A triplet at δ = 1.0 ppm
- **f** A septet at δ = 1.5 ppm
- **g** A quartet at δ =3.6 ppm

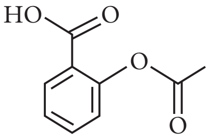
24 Consider the molecule hexane.

- **a** Draw all of the isomers with the same molecular formula as hexane.
- **b** Identify and compare the number of hydrogen and carbon environments in each isomer of hexane.

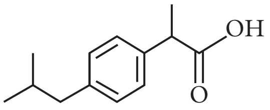
25 The structures of three drugs are shown below.



Paracetamol



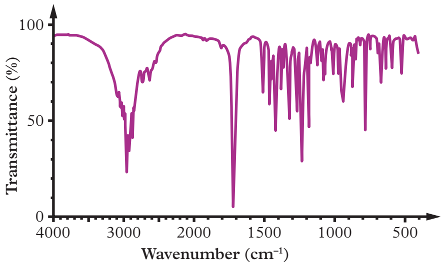
Aspirin



Ibuprofen

- **a** Identify the *m*/*z* of the molecular ion peaks of each of the drugs in a mass spectrum.

One of the three drugs was analysed using IR and the spectrum produced is shown below:



b Identify the functional groups in each of the drugs.

c Analyse the IR spectrum and identify the drug that was analysed using IR.

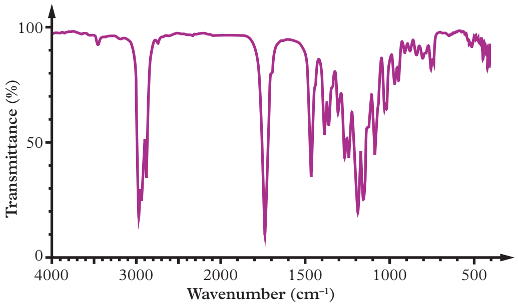
26 Consider the molecule octanol.

a Draw five structural isomers with the same molecular formula as octanol. One must have four methyl groups.

b Discuss whether IR would be useful to distinguish between the isomers you have drawn. Justify your response.

c Identify and compare the number of hydrogen and carbon environments in each isomer of octanol that you have drawn.

27 Analyse the following IR spectrum.

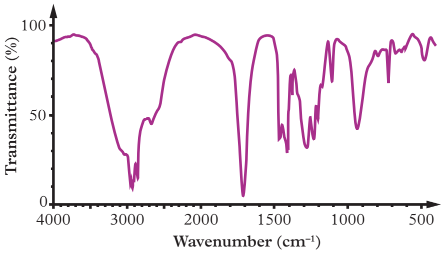


a Describe the features of the organic compound that produced this spectrum.

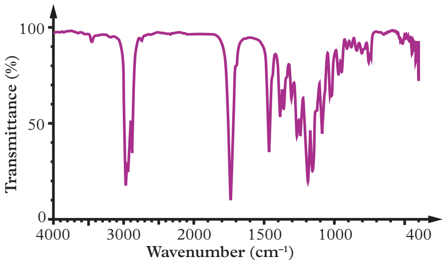
b Explain why it would be hard to identify an unknown compound that produced the IR with no other information.

28 Two molecules with the molecular formula $C_4H_8O_2$ are analysed using IR spectroscopy. Their spectra are shown below:

Spectrum 1



Spectrum 2



a Identify the functional group each IR spectrum is for and justify your reasoning.

b Determine and draw the molecule that produced spectrum 1.

c Based on the molecular formula, there are three isomers that could form spectrum 2. Name these isomers.

d Explain why it would not be easy to distinguish between the three isomers using a ^{13}C-NMR spectrum.

e The compound that formed spectrum 2 was analysed using ^{1}H-NMR and a spectrum with four peaks was produced. Identify the isomer that was analysed. Justify your answer.

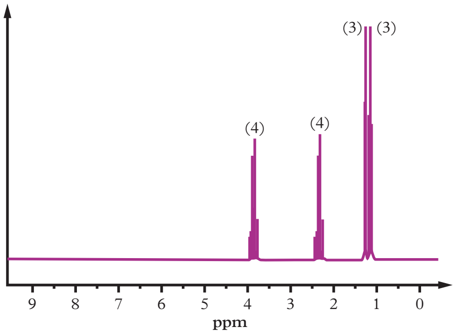
29 A molecule has carbon, hydrogen and one other atom in it. It produces a ^{13}C-NMR spectrum with three peaks and a ^{1}H-NMR spectrum with two peaks.

a Explain if you can determine its structural formula from only this information.

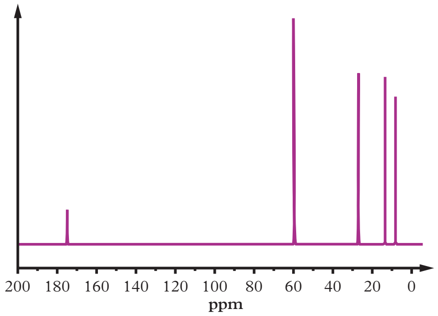
b Suggest what the molecule may be. Justify your response.

30 Determine the name and structure of the organic molecule that contains carbon, hydrogen and oxygen, has a molar mass of 102 g mol^{-1} and produces the following spectra:

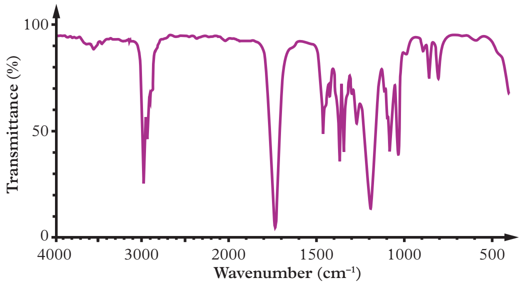
^{1}H-NMR spectrum



^{13}C-NMR spectrum

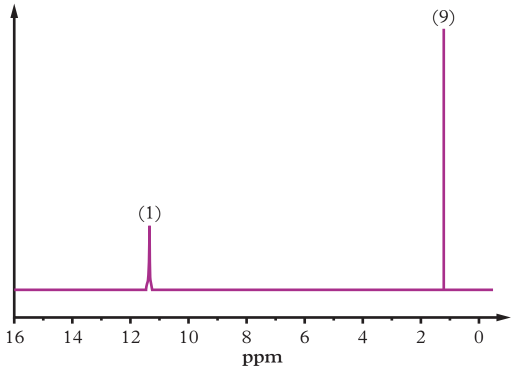


IR spectrum

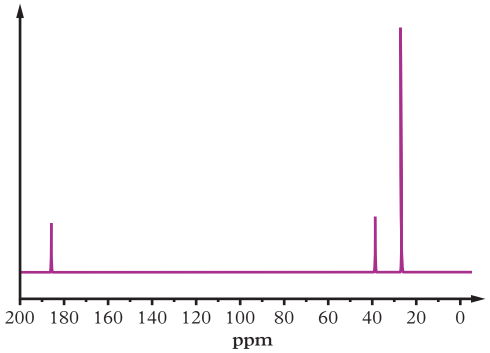


31 An isomer of the molecule in Question 30 produces the spectra below. Determine the structure and name of the organic compound.

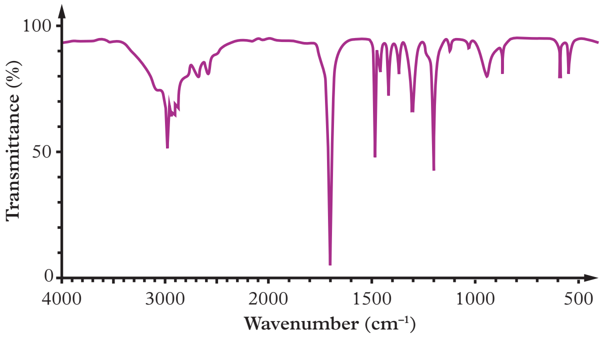
^{1}H-NMR spectrum



^{13}C-NMR spectrum

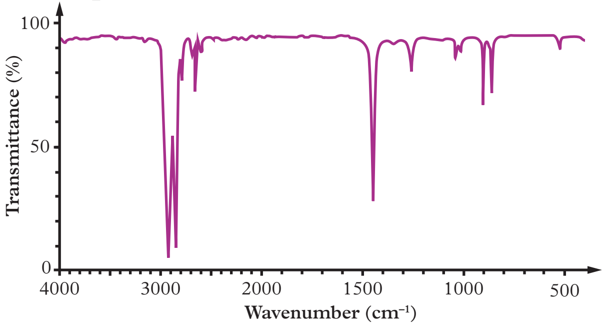


IR spectrum

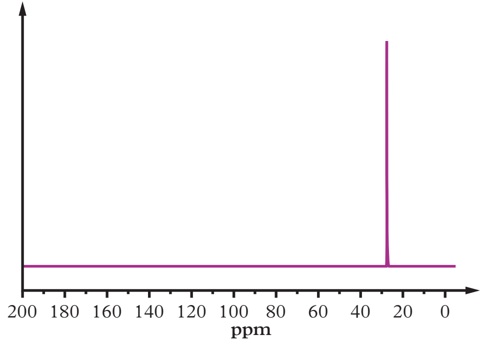


32 Determine the name and structure of the organic compound that contains carbon and hydrogen, has a molar mass of 84 g mol^{-1} and produces the spectra below:

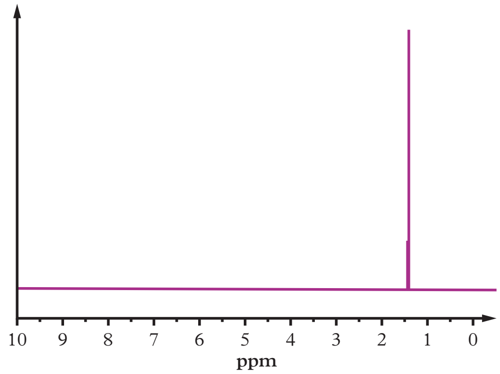
IR spectrum



^{13}C-NMR spectrum



^{1}H-NMR spectrum



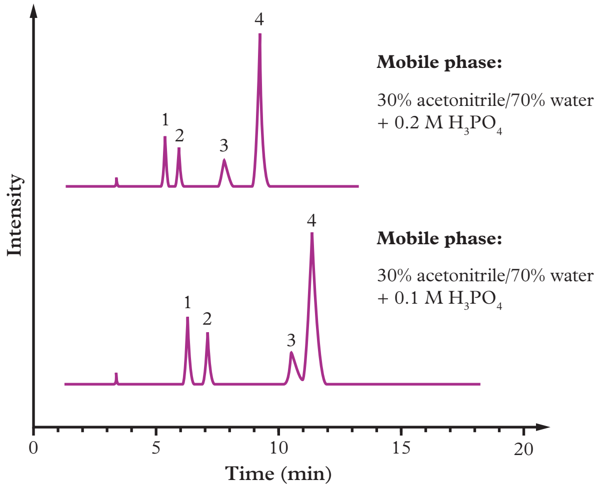
33 Determine the number of hydrogen and carbon environments in the following compounds:

- **a** $CH_3CH_2CH_2OCH(CH_3)_2$
- **b** $CH_3CH_2CH_2OCH_2CH_2CH_3$
- **c** $CH(CH_3)_2OCH(CH_3)_2$
- **d** $CH(CH_3)_2CHOCH_2CH_3$
- **e** $CH(CH_3)_2OCH(CH_3)_2$
- **f** $C(CH_3)_3CH_2OCH_3$

Design and discuss

34 Discuss how MS, IR, ^{1}H-NMR and ^{13}C-NMR can be used together to identify an unknown organic compound. Explain what unique information each provides.

35 The same sample is run through HPLC under the same conditions, except the mobile phase is changed for the second chromatogram.



1 - Serine (amino acid)
2 - Unknown compound
3 - Aminobutyric acid ($H_2N–(CH_2)_3–COOH$)
4 - Methionine

- **a** Describe the polarity of the stationary phase.
- **b** Explain why changing the concentration of the mobile phase would cause a difference in the retention times of the four compounds.
- **c** Discuss the polarity and structure of the unknown compound. Explain your reasoning.

36 Discuss whether you could analyse a water molecule using ^{1}H-NMR and ^{13}C-NMR. If so, explain what sort of spectra you would obtain.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER 12

Medicinal chemistry

KEY KNOWLEDGE

- → extraction and purification of natural plant compounds as possible active ingredients for medicines, using solvent extraction and distillation
- → identification of the structure and functional groups of organic molecules that are medicines
- → significance of isomers and the identification of chiral centres (carbon atom surrounded by four different groups) in the effectiveness of medicines
- → enzymes as protein-based catalysts in living systems: primary, secondary, tertiary and quaternary structures and changes in enzyme function in terms of structure and bonding as a result of increased temperature (denaturation), decreased temperature (lowered activity), or changes in pH (formation of zwitterions and denaturation)
- → medicines that function as competitive enzyme inhibitors: organic molecules that bind through lock-and-key mechanism to an active site preventing binding of the actual substrate.

Source: *VCE Chemistry Study Design (2023–2027)* reproduced by permission © VCAA

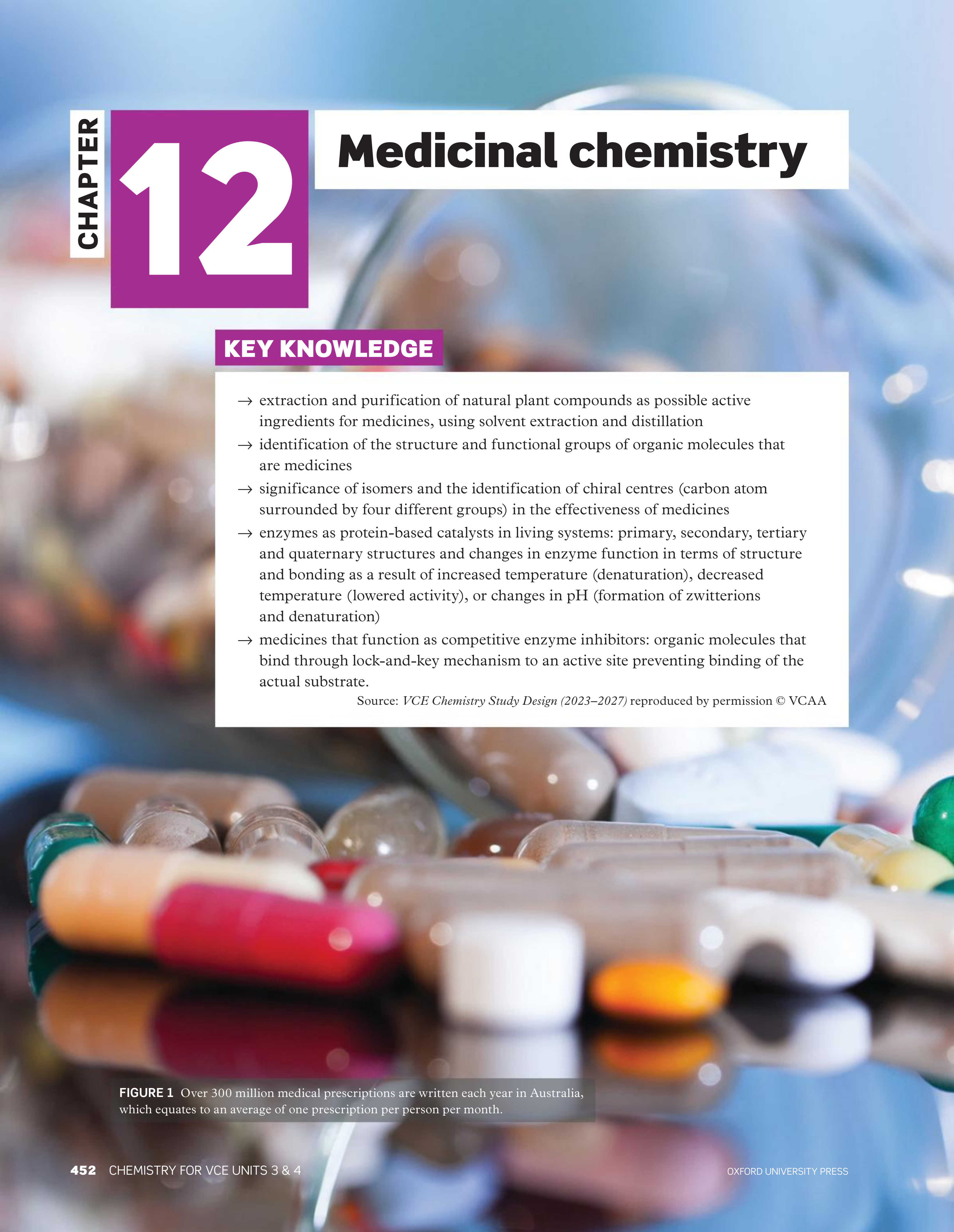


FIGURE 1 Over 300 million medical prescriptions are written each year in Australia, which equates to an average of one prescription per person per month.

GROUNDWORK

In Chapter 12, you will learn about organic molecules in medicinal chemistry, including the importance of structure in their effectiveness. You will also explore enzymes and how medicines can target enzymes to modify their function.

This chapter will build on concepts you have already learnt in Chapters 5 and 8. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

12A Explain what a catalyst is.

12A Groundwork resource
Catalysts

12B Distinguish between an aldehyde and a ketone.

12B Groundwork resource
Functional groups

12C Compare the strengths of the three types of intermolecular forces.

12C Groundwork resource
Intermolecular forces

PRACTICALS

12.1 **PRACTICAL: FIELDWORK**	How can you extract active ingredients from medicinal plants?	pro
12.3 **PRACTICAL: CLASSIFICATION & IDENTIFICATION**	Can you identify chiral centres in popular medicinal compounds?	Page 524
12.4 **PRACTICAL: MODELLING**	How can paper be used to model a simple protein?	Page 526
12.5 **PRACTICAL: LITERATURE REVIEW**	Does the lock-and-key model fit?	pro

12.1

Extracting and purifying plant compounds for medicines

KEY IDEAS

In this topic, you will learn that:

- plants contain many bioactive compounds that can affect human health
- medicinal compounds in plants can be extracted and purified by solvent extraction or distillation.

So far, you have learnt about the many uses of organic compounds, including in medicines and other pharmaceutical products. This is because some organic compounds are **bioactive**. In other words, they are **active ingredients** that can bind to **biological targets** in the human body and affect one or more chemical processes.

These active ingredients can be synthesised using, for example, the reactions you learnt about in Chapter 9. Others are sourced directly from nature. In this topic, you will learn about medicines extracted directly from plant material.

bioactive
can bind to targets and affect chemical processes in the human body

active ingredient
a compound that is bioactive; in medicine, it is the compound that has medicinal properties

biological target
a target (e.g. a receptor or enzyme) in a living system (e.g. the human body) to which a compound can bind and exert a biological effect

secondary metabolite
a chemical substance produced by a plant that protects it from predators (e.g. herbivores); typically does not directly affect the growth of the plant

Medicines from plants

Plants provide nutrition for primary consumers in the food web. Because they are fixed into the ground and cannot escape herbivores, plants have evolved complex chemical defences against predators to prevent themselves from being eaten. The compounds produced are called **secondary metabolites**.

These compounds are either bitter-tasting, poisonous or both. Examples include caffeine in coffee, catechins in tea and nicotine in tobacco plants, which deter herbivores from eating their leaves.

Some of the compounds produced by plants for their own self-defence also have unrelated actions in the human body. Beta-carotene, which plants produce to reflect harmful ultraviolet rays, is required for healthy human vision. Caffeine in coffee and theobromine in cocoa deter herbivores, but stimulate the human central nervous system to make us feel more alert. Isoflavones encourage soy plant root cells to interact with helpful bacteria in the soil, but the same isoflavones have dual effects as both an antioxidant and as an oestrogen mimic in the human body.

Tens of thousands of plant compounds have been identified, extracted, purified and tested. Many are now used as modern medicines.

FIGURE 1 Tea growers on a tea farm in southern China use goats to control weed growth. Goats eat weeds but do not eat the tea plants because the catechins and caffeine in tea leaves make them unappetisingly bitter for goats.

12.1

PRACTICAL: FIELDWORK

How can you extract active ingredients from medicinal plants?
Go to your **Student obook pro**.

Artemisinin to treat malaria

Blue-green wormwood is listed as a treatment for fever, inflammation and malaria in an ancient Chinese text dating back to 340 BCE. Despite being used for thousands of years as a traditional herbal medicine, it did not undergo rigorous scientific testing until the late 1960s.

At this time, Professor Tu Youyou identified, extracted, purified and tested a particular blue-green wormwood extract called artemisinin. She was awarded a Nobel Prize in Medicine and Physiology for her discovery that artemisinin acts as a powerful treatment for multidrug-resistant malaria.

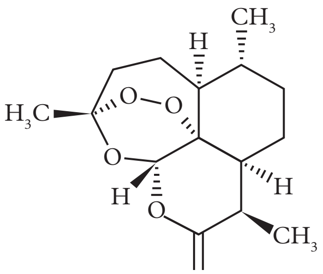


FIGURE 2 Artemisinin ($C_{15}H_{22}O_5$)

Artemisinin contains an unusual peroxide bridge (–O–O–), which is crucial to its activity as an antimalarial drug (Figure 2). You will learn more about the role of functional groups in the effectiveness of medicines in Topic 12.2.

FIGURE 3 Blue-green wormwood plants

FIGURE 4 Opium poppies growing in Afghanistan

FIGURE 5 Jaborandi plant native to South America

Morphine from opium poppies

Morphine is a secondary metabolite that deters herbivores from opium poppies and also treats severe pain in humans. This molecule (Figure 6) was isolated and purified between 1803 and 1805 by German pharmacist Friedrich Sertürner. By the early 19th century, injectable morphine was widely available as a painkiller. Today, morphine is on the WHO Model List of Essential Medicines.

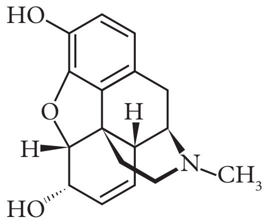


FIGURE 6 Morphine ($C_{17}H_{19}NO_3$)

Pilocarpine from jaborandi

Pilocarpine is a bitter-tasting substance that acts as a chemical defence for jaborandi plants. In humans, pilocarpine is used to treat glaucoma due to its ability to reduce pressure inside the eye by restricting the constriction of the pupil after its dilation.

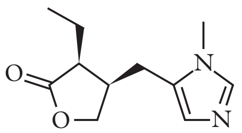


FIGURE 7 Pilocarpine ($C_{11}H_{16}N_2O_2$)

Extracting active molecules from plants

Plants contain thousands of compounds, only some of which might be bioactive. Plant material must therefore be purified by solvent extraction or distillation to obtain the active ingredient. Both of these methods require a thorough understanding of the properties of the desired compound, as well as the properties of the other compounds it is being separated from.

Solvent extraction

Solvents can be used to extract active ingredients from plants. A solvent is chosen that dissolves only the desired compound, separating it from the other components of the plant. Polar solvents (such as water or ethanol) or non-polar solvents (such as hexane or trichloromethane) are selected depending on the structure and polarity of the compound. This process is called **solvent extraction** (Figure 8). The dissolved compound of interest can be filtered to separate it from the remaining plant material.

solvent extraction
a process by which a desired compound is separated from a mixture because it is soluble in a solvent, but the other components in the mixture are not soluble

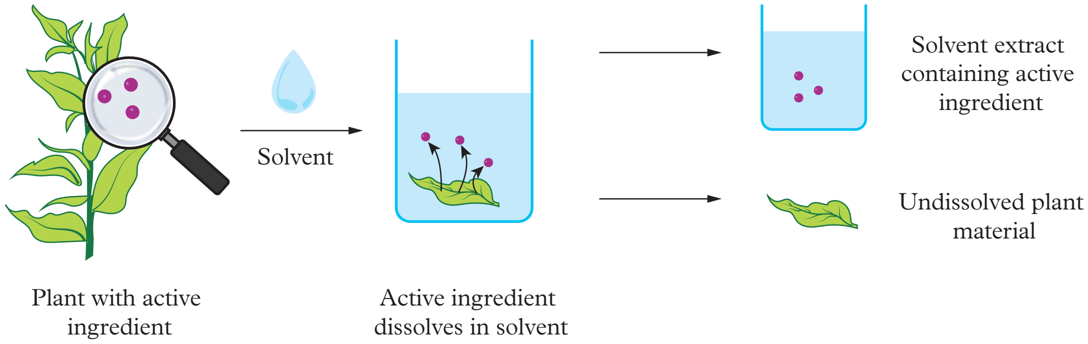


FIGURE 8 Active ingredients (purple) can be extracted from plant material using solvent extraction.

Distillation

Distillation can separate the compound of interest from the rest of the mixture by boiling point. The process is described in the steps below:

1. The plant is placed in a solvent mixture.
2. The plant–solvent mixture is heated at the bottom of a vertical distillation chamber or column.
3. As the mixture is heated and the components in the plant reach their boiling point, they will evaporate.
4. Vapour rises to the top of the chamber, then cools down as it travels through a condenser and reforms as a liquid.
5. The liquid is collected in a container (e.g. a beaker).
6. Compounds with the lowest boiling points will vaporise at low temperatures and can be extracted first. Compounds with the highest boiling point will be extracted last.

Ideally, the compound of interest will have a very different boiling point from the other components in the plant–solvent mixture. This will allow you to extract a pure **fraction**. Unfortunately, this is rarely the case. Instead, the process can be repeated to improve purity. You can also reduce the increments for each temperature range (e.g. use 100–105°C and 105–110°C instead of 100–110°C) to capture your desired compound.

distillation
a process by which a desired compound is separated from a mixture because it has a different boiling point to the other components in the mixture

fraction
a substance or mixture of substances with the same or a similar boiling point

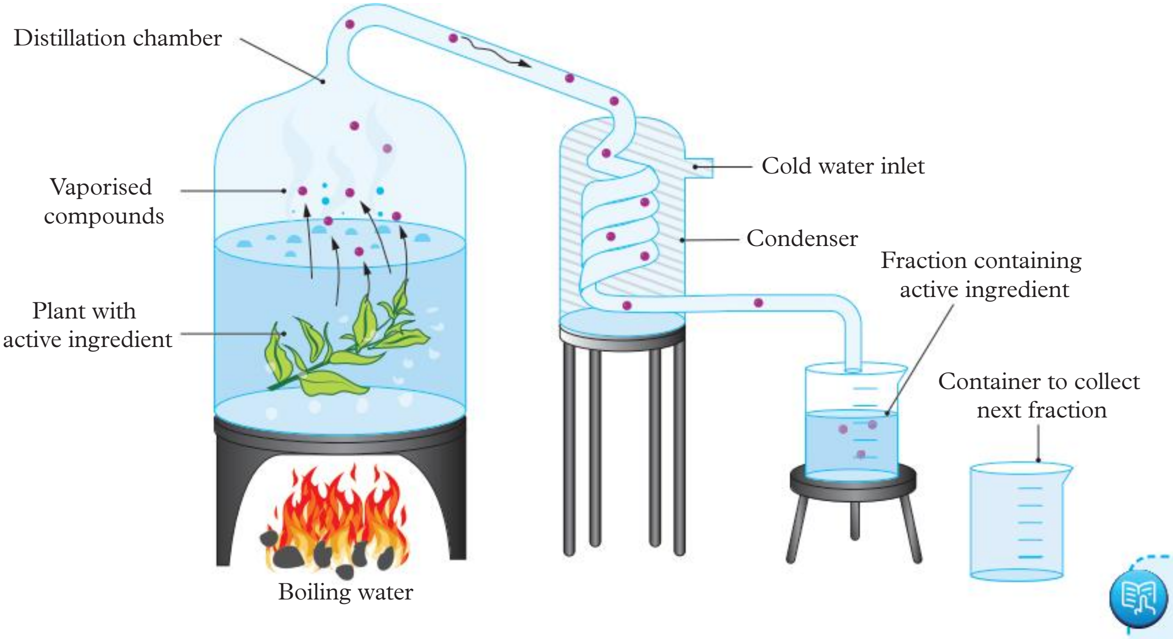


FIGURE 9 Active ingredients (purple) can be extracted from plant material using distillation.

12.1A Challenge Find me in your obook pro

12.1B Challenge Find me in your obook pro

If you are up for something more difficult, try comparing synthetic and natural medicines in Challenge 12.1A, or investigate the medicinal properties of native Australian plants in Challenge 12.1B.

12.1 CHECK YOUR LEARNING

Describe and explain

1 Explain how medicines can be derived from plants. Use the terms *bioactive* and *active ingredient* in your response.

2 Describe how tea and coffee affect the human body.

3 Explain what a secondary metabolite is. Use an example to explain how they are used by plants.

Apply, analyse and compare

4 Compare the role of pilocarpine in jaborandi plants with how pilocarpine is used medicinally.

5 Compare the processes involved in distillation and solvent extraction.

  a Identify the physical property of the active ingredient that must be known to carry out each method of separation.

  b Suggest a scenario where one method may be preferred over the other.

Design and discuss

6 Patchoulol from the patchouli plant is commonly used in fragrances, and also has anti-inflammatory, anti-influenza and anti-cancer effects. The compound has a boiling point of 280–290°C, is insoluble in water but soluble in ethanol and diethyl ether. It is typically extracted from the patchouli plant using distillation, but the patchoulol extract has a low purity (~40%).

  a Evaluate the use of distillation and suggest reasons for the low purity.

  b Discuss a strategy the chemist could use to improve the purity of patchouli extracted.

7 Design a flowchart to outline the four steps for the approval of a new medicine. Include details about the three stages of clinical trials in your flowchart.

12.2

Structure and functional groups of medicines

KEY IDEAS

In this topic, you will learn that:

- organic molecules that are medicines can contain many different functional groups.

Most organic molecules that are medicines have more complex structures compared to those you studied in Chapter 8. Many contain multiple functional groups per molecule. The diversity of medicinal compounds available is largely a result of the vast number of ways in which these functional groups can be combined.

Study tip

Go back to Chapter 8 if you need to refresh your understanding of functional groups, and Chapter 9 to review how functional groups affect the chemical reactions that molecules can participate in.

Functional groups in medicinal compounds

Remember that functional groups are atoms or groups of atoms that are responsible for a molecule's ability to undergo specific chemical reactions. The functional groups present and the way in which they are arranged in a molecule are very important. This helps the molecule bind to a desired biological target. Therefore, the shape of the molecule gives it its medicinal properties.

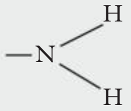
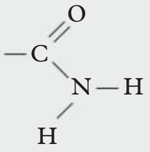
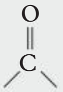
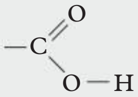
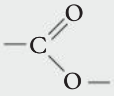
The functional groups you learnt previously in VCE Chemistry are summarised in Table 1.

Worked example 12.2 walks you through the process of identifying functional groups in a medicinal organic molecule.

Study tip

Although you do not need to know how to name ethers, it is handy to be able to identify them in a chemical structure.

TABLE 1 Functional groups found in organic molecules

Functional group	Structure
alkane	C–C
alkene	C=C
haloalkane	–F, –Cl, –Br or –I
primary amine	
primary amide	
hydroxyl	$-O-H$
carbonyl	
carboxyl	
ester	
ether	$-O-$

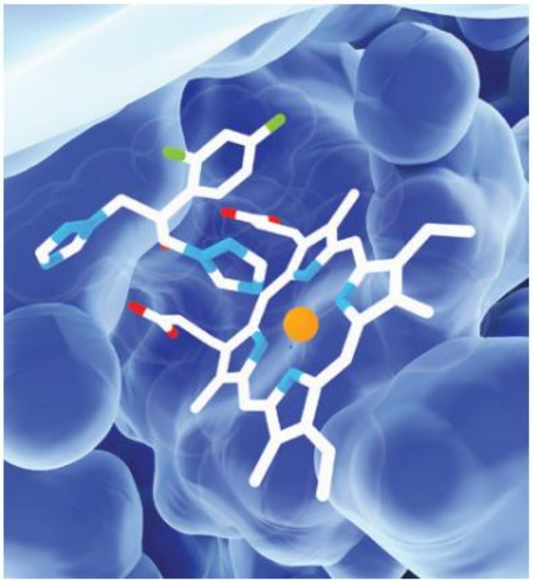


FIGURE 1 Functional groups allow a medicinal compound, like fluconazole, to bind to its biological targets. Here, it binds to the lanosterol 14 alpha-demethylase enzyme.

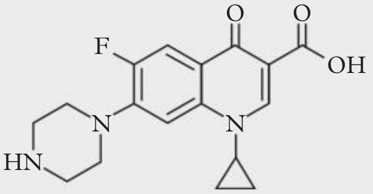
12.2 Worked example
Find me in your obook pro

12.2 Worked example
Video demonstration

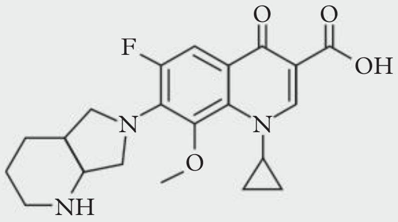
12.2 REAL-WORLD CHEMISTRY

A class full of antibiotics

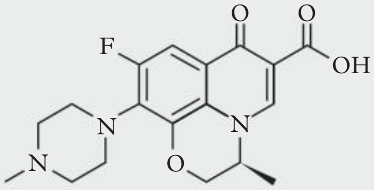
Medicinal compounds are often divided into classes depending on their chemical structure. The three structures below belong to a class called quinolones. They are used as antibiotics because of their ability to kill bacteria.



Ciprofloxacin



Moxifloxacin



Levofloxacin

Apply your understanding

1 Identify the functional groups in each structure.

2 Compare the three structures and predict which functional groups give the compounds their medicinal properties. Hint: They bind to the same biological target.

3 Suggest a reason why these medicines are named similarly (with the suffix –*floxacin*).

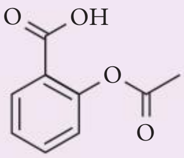
12.2 CHECK YOUR LEARNING

Describe and explain

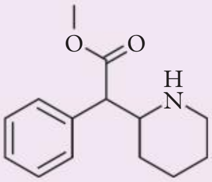
1 Explain why the structure and functional groups are important for the medicinal properties of an organic compound.

Apply, analyse and compare

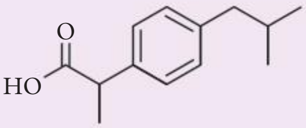
2 Aspirin is a type of pain relief medication. Circle and label two functional groups on the structure of aspirin.



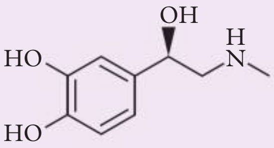
3 Methylphenidate is used to treat attention deficit hyperactivity disorder (ADHD). Circle and label two functional groups on the structure of methylphenidate.



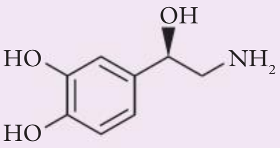
4 Ibuprofen is an anti-inflammatory pain relief medication. Circle and label one functional group on the structure of ibuprofen.



5 Adrenaline and noradrenaline are very important compounds involved in regulating the nervous system in the human body. They are structurally and functionally similar. The structures of both are shown below.



Adrenaline



Noradrenaline

a Identify the functional groups in adrenaline and noradrenaline.

b Compare the two structures and predict which functional groups make the compounds bioactive. Hint: They bind to the same biological target.

12.3

Chirality in medicine

KEY IDEAS

In this topic, you will learn that:

- stereoisomers are molecules with at least one chiral carbon atom
- stereoisomers of medicinal compounds can have very different functions in the body because of their ability to bind biological targets.

In Chapter 8, you learnt about structural isomers of organic compounds. These compounds have the same molecular formula but their atoms are connected in different ways. Organic compounds can also have **stereoisomers**. Stereoisomers contain atoms that are bonded in the same way. However, they have different three-dimensional configurations. This affects the way that they interact with biological targets.

stereoisomers
molecules with the same molecular formula that contain atoms connected in the same way, but with a different three-dimensional configuration

Stereoisomerism

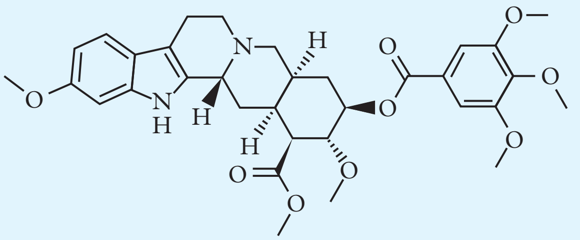
Stereoisomerism exists due to the presence of one or more **chiral centres** in a compound. In VCE Chemistry, a chiral centre is any carbon atom with single bonds to four different atoms or groups of atoms. They are also called chiral carbons. When the bonds around a chiral centre rotate, they result in different three-dimensional configurations. Chiral carbons are typically indicated in a structural formula using asterisks (*). Worked example 12.3A shows you how to identify chiral centres.

chiral centre
a carbon atom with single covalent bonds to four different atoms or groups of atoms

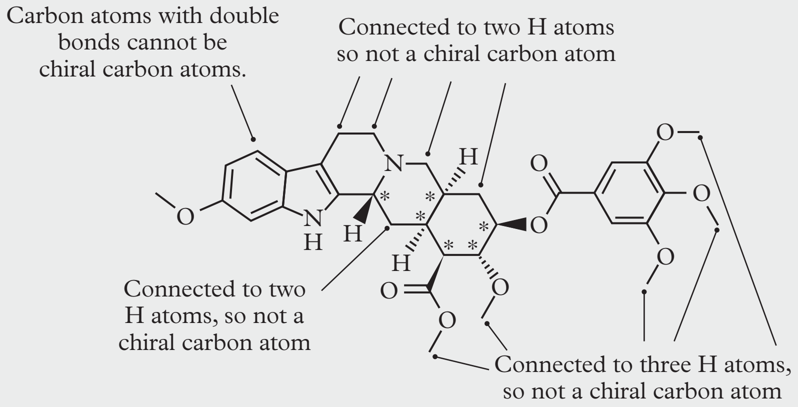
12.3A WORKED EXAMPLE

IDENTIFYING CHIRAL CENTRES

Reserpine is a medication extracted from a plant called Indian snakeroot. It is used to treat both high blood pressure and psychosis. Use asterisks to show the chiral centres in this molecule.



Solution

Think	Do
Remember that in a structural formula, the vertices represent carbon atoms. Look at these carbon atoms and identify the atoms (or groups of atoms) that they are bonded to. If you can count four different atoms (or groups of atoms) around the carbon atom, it is a chiral centre. Use asterisks to show these.	 There are six chiral centres. All of these carbon atoms are connected to four different atoms or groups of atoms.

12.3
PRACTICAL: CLASSIFICATION & IDENTIFICATION

Can you identify chiral centres in popular medicinal compounds? Go to page **524**.

You will notice that the three-dimensional arrangement of atoms around a chiral carbon atom is shown using wedged and dashed lines, instead of the typical thin lines. The wedged line indicates that the bond is coming out of the page, towards you. The dashed line indicates that the bond is going into the page, away from you.

An example of a pair of stereoisomers, each drawn in two different ways, is shown in Figure 1 below. The first structure of each pair is drawn with dashed and wedged lines. You can also see this in Worked example 12.3B.

12.3B Worked example
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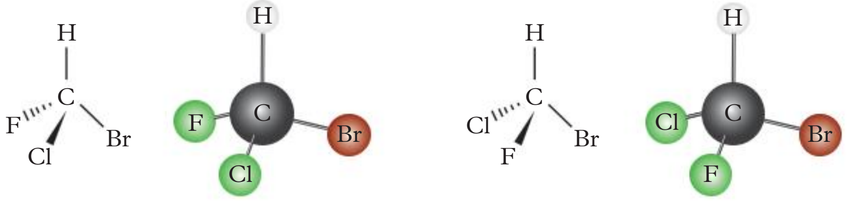
12.3B Worked example
Video demonstration



(S)-bromochlorofluoromethane **(R)-bromochlorofluoromethane**

FIGURE 1 The two stereoisomers of bromochlorofluoromethane, each with one chiral carbon

Study tip

You may come across the term *optical isomer*. This is a form of stereoisomer. In VCE Chemistry, we can treat them as the same thing.

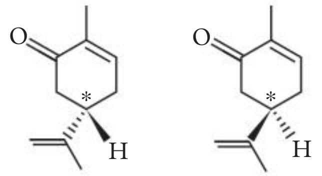
Stereoisomers are named using various prefixes: (S) or (R), L- or D-, (+) or (–). In VCE Chemistry, you do not need to know how to use the prefixes, but they have been provided in this topic just so you can tell the stereoisomers apart.

Stereoisomers in nature

The difference in three-dimensional configuration of stereoisomers allows organic compounds to interact with other chemicals in different ways. One example in nature is carvone. Carvone has one chiral centre, indicated by the asterisk in Figure 2. The two stereoisomers of carvone have different aromas because they bind differently to the receptors in our body that perceive taste and smell.

(S)-carvone has a caraway aroma, making it useful in potato salads. (R)-carvone is present in dill, giving it a minty aroma for use in tzatziki.

The two stereoisomers of carvone also bind differently to chemicals in other organisms. (R)-carvone (found in mint or dill) is used as a mosquito repellent because of the way it can interact with biological targets in the mosquito. (S)-carvone (found in caraway) binds to target compounds in potatoes to prevent premature sprouting and therefore prolong the storage time of the potatoes.



(S)-carvone **(R)-carvone**

FIGURE 2 (S)-carvone has a caraway aroma and (R)-carvone has a minty aroma.

Stereoisomers in medicine

Stereoisomerism is also very important in medicinal compounds. Often, only one of the stereoisomers binds well to biological targets in the human body (e.g. receptors or enzymes). The other one may bind poorly, not bind at all, or may even be toxic.

If the other stereoisomer is harmless, it will often be included in the finished product because it can be very expensive to remove. Mixtures containing an equal amount of two opposing stereoisomers are called **racemic** (*ra-see-mik*) **mixtures**. Have a look at Skill drill 12.3 to evaluate the use of racemic mixtures for medicines.

racemic mixture
a mixture that contains an equal concentration of stereoisomer pairs

Three medicines that are sold as racemic mixtures are: ibuprofen, methadone and synthetic vitamin C. We will explore how the stereoisomers of these three medicinal compounds influence their effectiveness as medicines.

12.3 Skill drill
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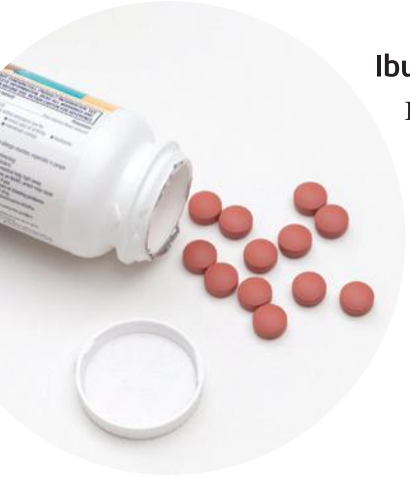
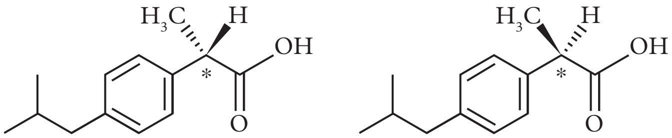


FIGURE 3 Ibuprofen tablets

Ibuprofen – stereoisomerism affects effectiveness

Ibuprofen is a common, anti-inflammatory pain relief medicine used widely in Australia. It works by decreasing the production of inflammatory molecules (especially prostaglandins) in the body. Ibuprofen has one chiral centre (Figure 4). The (S) stereoisomer is approximately 200 times more effective than the (R) stereoisomer because it binds better to its biological target, the cyclooxygenase enzyme.

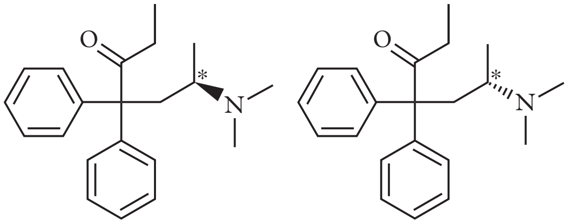


(S)-ibuprofen **(R)-ibuprofen**

FIGURE 4 (S)-ibuprofen is more effective than (R)-ibuprofen.

Methadone – stereoisomerism affects target binding

Methadone is a powerful opioid pain relief medication. It has one chiral centre (Figure 5). The (R) stereoisomer binds 20 times better to the opioid receptor in the body than the (S) stereoisomer.



(S)-methadone **(R)-methadone**

FIGURE 5 (R)-methadone is much more effective than (S)-methadone.

Study tip

There are several ways to name stereoisomers, including (R) and (S), D- and L- and (+) and (–). You do not need to learn these naming systems in VCE Chemistry.

Synthetic vitamin C – stereoisomerism affects bioactivity

Vitamin C (also known as ascorbic acid or E300) is a sour-tasting compound found in citrus fruits. It has two chiral centres (Figure 7). In nature, vitamin C is found in the L-ascorbic acid stereoisomer form. In animals, this is a powerful antioxidant that also plays an important role in the growth and repair of body tissues. Most animals can synthesise their own vitamin C, but some (including humans) are unable to do so and must obtain vitamin C from their diet.

Synthetic vitamin C supplements are available widely in Australia. They are made by the chemical conversion of glucose into a racemic mixture of vitamin C. The other D-ascorbic acid stereoisomer is not biologically active. However, as it is harmless, the product can be sold as a racemic mixture.

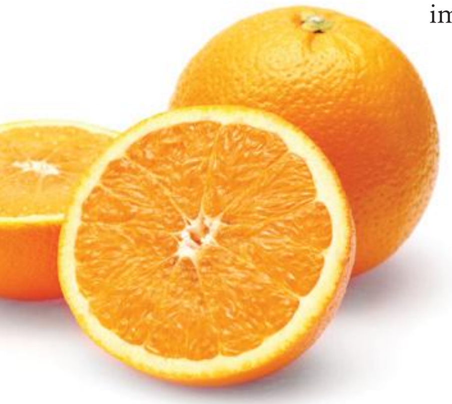
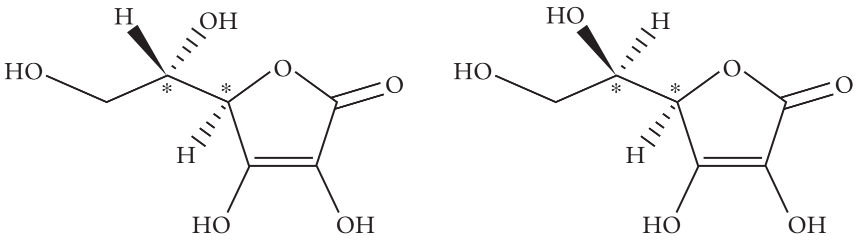


FIGURE 6 Ascorbic acid is found naturally in citrus fruits.



L-ascorbic acid **D-ascorbic acid**

FIGURE 7 L-ascorbic acid is biologically active and D-ascorbic acid is not.

12.3 CHECK YOUR LEARNING

Describe and explain

1 Explain what a chiral centre is.

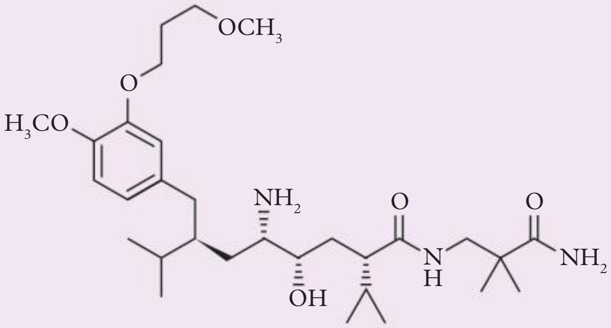
2 Explain what a stereoisomer is and how it is different from a structural isomer.

3 Explain why it is important to consider stereoisomerism of organic compounds that are used as medicines.

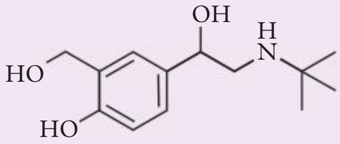
Apply, analyse and compare

4 Identify all the chiral centres in the medicines below by drawing an asterisk (*) where appropriate.

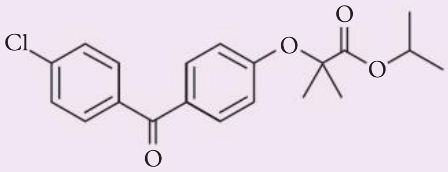
a Aliskiren is a medication for high blood pressure.



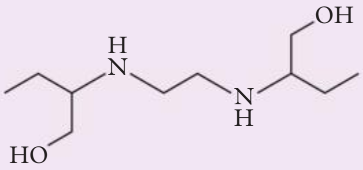
b Salbutamol is the active ingredient in asthma puffers.



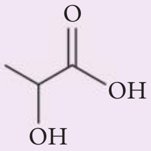
c Fenofibrate is an antihistamine used to treat allergies like hay fever.



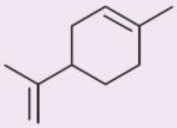
d Ethambutol is used to treat tuberculosis.



5 Lactic acid is shown below. Draw the structural formula for a stereoisomer of lactic acid with dashed and wedged lines.



6 Limonene is a terpene molecule found naturally in plants. Limonene has one chiral carbon atom, which results in two stereoisomers. D-limonene smells like citrus fruits, while L-limonene smells like pine trees.



a Circle the chiral centre in limonene.

b State two properties you would expect to be the same in the two stereoisomers of limonene.

c State two properties you would expect to be different in the two stereoisomers of limonene.

12.4

Enzymes as protein-based catalysts

KEY IDEAS

In this topic, you will learn that:

- enzymes are proteins that catalyse specific chemical reactions in living cells
- reactants (substrates) bind to enzymes via a specific region called the active site
- the interactions between an enzyme and substrate are described using the lock-and-key model
- the three-dimensional structure of the active site is determined by the primary, secondary, tertiary and quaternary structure of the enzyme
- enzymes are sensitive to changes in temperature and pH, which can alter their effectiveness as biological catalysts.

Previously, you learnt about catalysts and their role in increasing the rate of chemical reactions. One class of catalysts is biological catalysts or **enzymes**. These are almost always proteins that are manufactured in the body to speed up reactions in living cells. Enzymes are very common biological targets for medicines.

enzyme
a biological, protein-based catalyst that speeds up chemical reactions

Enzymes are biological catalysts

Enzymes work by binding to one or more target molecules or **substrates**. This binding occurs at a specific region of the enzyme called the **active site**. Binding results in the formation of an **enzyme–substrate complex**. The specific catalyst function of the enzyme varies depending on what it is, but examples include:

- bringing together two or more substrates
- breaking bonds in a single substrate so that it can then participate in another chemical reaction.

The products of the catalysis are then released from the enzyme and it is free to bind new substrates.

substrate
a reactant in an enzyme-catalysed reaction

active site
the region of an enzyme that binds one or more substrates

enzyme–substrate complex
the structure formed when one or more substrates bind to the active site of an enzyme

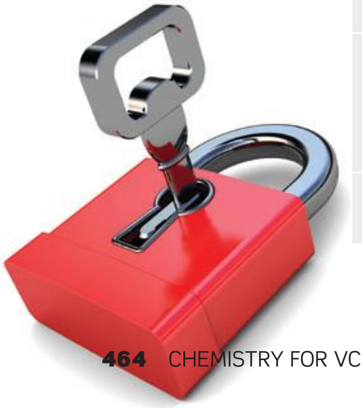
Lock-and-key mechanism

Because the substrate(s) fit into the enzyme similarly to the way a key fits into a lock, this process is often described as the lock-and-key mechanism. You can see how they are similar in Table 1.

TABLE 1 Similarities between locks and keys and the mechanism of enzyme action

Locks and keys	Enzyme
Only one particular type of key fits into the lock.	Only one particular type of substrate (or combination of substrates) fits into the active site.
When the key is inserted, it connects with the inside of the lock and causes small changes within the lock.	Once the substrate has bound to the active site, it forms temporary bonds with the enzyme that causes small changes in the structure of the substrate.
It is almost impossible to open the lock without the key.	The rate of reaction is usually almost zero in the absence of the enzyme.

FIGURE 1 A lock and key



12.4
PRACTICAL: MODELLING
How can paper be used to model a simple protein?
Go to page **526**.

The shapes of the enzyme and substrate(s) are very important. In the lock-and-key model, they must be complementary (fit perfectly with each other). You can see this in Figure 2.

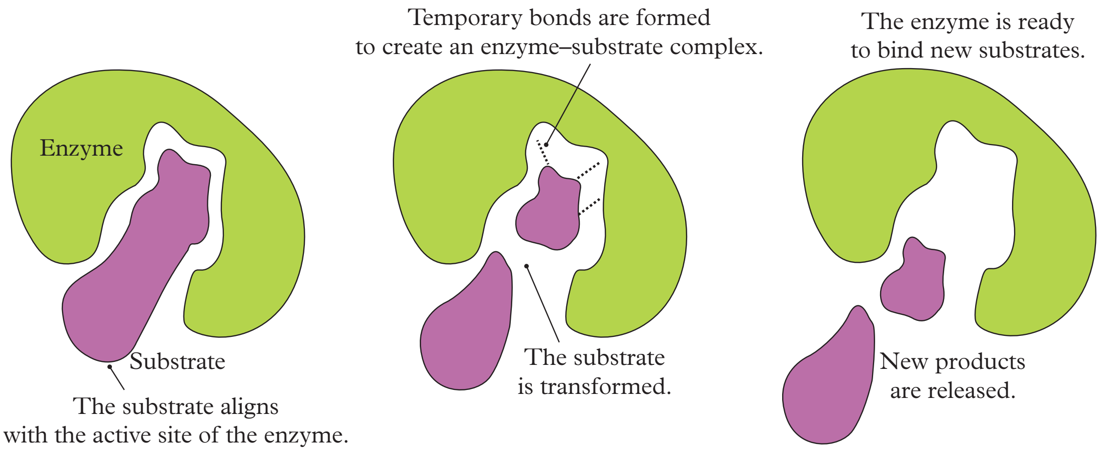


FIGURE 2 An example of enzyme action

The structure of enzymes

It is very clear that the three-dimensional shape of the enzyme, particularly at its active site, is very important for its ability to bind substrates and therefore function as a catalyst. There are four levels of organisation:

- primary structure
- secondary structure
- tertiary structure
- quaternary structure.

Primary structure of enzymes

Enzymes are proteins, meaning that they are composed of chemicals called amino acids. The **primary structure** of enzymes refers to the sequence of amino acids linked together by covalent amide bonds or peptide linkages. This forms a polypeptide (*poly* = many, *peptide* = amino acid) chain. You learnt about the reactions that form polypeptides in Chapter 9.

Secondary structure of enzymes

The **secondary structure** of enzymes is formed by the hydrogen bonding between C=O of an amide group and N–H of a nearby amide group. This allows the polypeptide chain to become organised into two new shapes:

- an **alpha helix** coil, or
- a folded **beta-pleated sheet**.

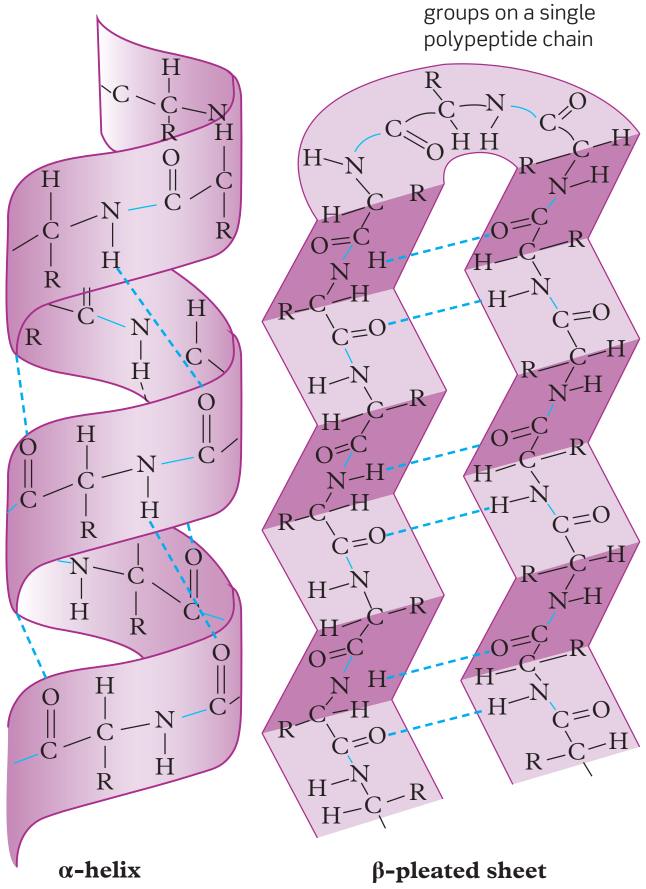


FIGURE 3 The polypeptide chain can be further organised into an alpha helix or a beta-pleated sheet.

primary structure
the structure formed when amino acids are joined in a sequence by covalent amide bonds or peptide linkages

secondary structure
the structure formed when amide groups along the polypeptide chain form hydrogen bonds with each other; forms an alpha helix or beta-pleated sheet

alpha helix
a coiled structure formed from hydrogen bonding between the amide groups on a single polypeptide chain

beta-pleated sheet
a folded structure formed from hydrogen bonding between the amide groups on a single polypeptide chain

Tertiary structure of enzymes

tertiary structure the structure formed when covalent bonds, hydrogen bonds, ionic interactions or dispersion forces form between side groups of amino acids along a single polypeptide chain; forms the three-dimensional structure of an enzyme

The **tertiary structure** of enzymes gives them their three-dimensional shape. This forms from covalent bonds, hydrogen bonds, ionic interactions and dispersion forces between side groups on amino acids along a single polypeptide chain.

The types of interactions holding together the tertiary structure depends on the specific amino acids present, and the functional groups they contain. For example, if a side group can become ionised, it can form ionic interactions with other charged side groups. Side groups with acid–base properties (i.e. they can donate or accept protons) can gain a charge, allowing them to form ionic interactions. Similarly, sulfur (–SH) groups can undergo oxidation to form a covalent bond called disulfide bridge (S–S bond).

The different types of bonds involved in the tertiary structure of enzymes are summarised in Figure 4 and Table 2.

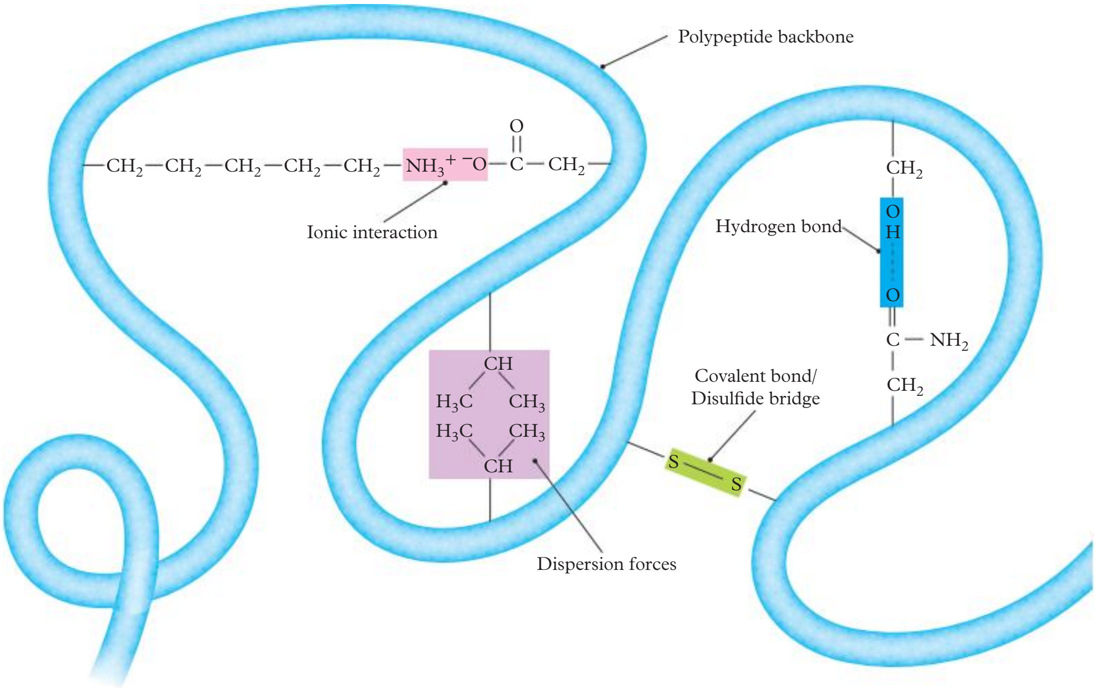


FIGURE 4 A tertiary structure forms interactions between amino acid side groups.

TABLE 2 Interactions between amino acid side groups in a tertiary structure

Interaction	Occurs between...	Example amino acids
Covalent bonds	sulfur-containing side groups; also specifically called disulfide bridges	cysteine
Hydrogen bonds	polar side groups, where a hydrogen atom is covalently bonded to a fluorine, oxygen or nitrogen atom	asparagine, glutamine
Ionic interactions	an acidic and a basic side group	aspartic acid, arginine
Dispersion forces	non-polar side groups	alanine, glycine

Quaternary structure of enzymes

The **quaternary structure** is formed via the same types of bonding as the tertiary structure. However, the bonding occurs between different polypeptide chains. Haemoglobin is one example of a protein that consists of four polypeptide chains bonded together.

The four levels of enzyme structure are summarised in Figure 5.

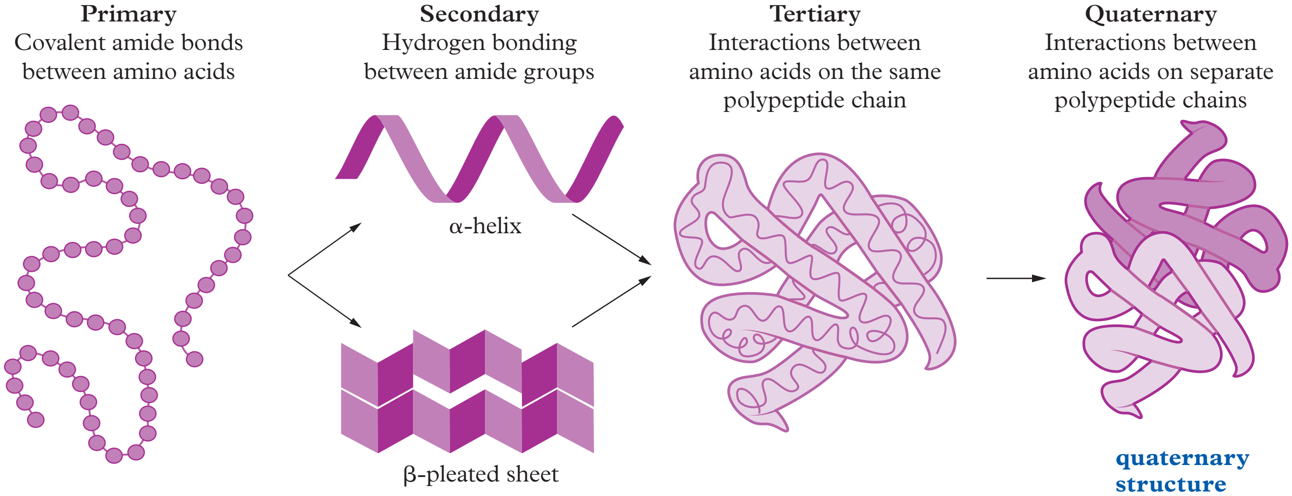


FIGURE 5 Summary of the four levels of enzyme structure

quaternary structure
the structure formed when covalent bonds, hydrogen bonds, ionic interactions or dispersion forces form between side groups of amino acids of two or more polypeptide chains

Factors affecting enzyme activity

Each enzyme operates at an optimal temperature and pH. This is where their catalytic activity is at a maximum. For example, trypsin, which digests proteins in the small intestine, works best at a temperature of 40°C and a pH of 8. Taq polymerase, which catalyses the formation of new DNA strands in a microorganism called *thermus aquaticus*, functions optimally at 75°C and pH 9. Outside of these conditions, the enzyme's ability to catalyse different reactions can be reduced.

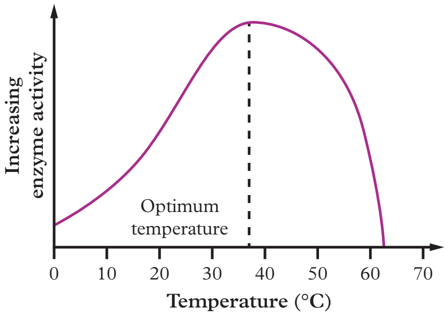


FIGURE 6 Effect of temperature on enzyme activity

Effect of temperature on enzyme activity

Most enzymes in the human body work best around 37°C (Figure 6).

At low temperatures, enzyme activity is lowered

Remember from Chapter 5 that decreasing temperature decreases the rate of a chemical reaction. The same applies for chemical reactions catalysed by enzymes.

When temperature is decreased:

- the substrate particles vibrate more slowly and the frequency of successful collisions with the active site of the enzyme is reduced
- there may not be enough energy in the system to overcome the activation energy for the catalysis to occur.

This reduction in the enzyme activity is represented by the decreasing slope in the curve of Figure 6 (from 37°C to 0°C). At very low temperatures, the enzyme can become inactive. However, this change is reversible, since increasing the temperature will restore enzyme activity.

Study tip

Not all enzymes have an optimal temperature of 37°C, but many enzymes found in warm-blooded animals, such as humans, do.

At high temperatures, enzymes become denatured

Increasing temperature increases the rate of a chemical reaction. However, it can destabilise the bonds that hold the enzyme in its secondary, tertiary, and quaternary structures. This process is called **denaturation**.

denaturation
the process by which an enzyme loses its three-dimensional structure and can no longer bind to its natural substrates

The folding of proteins like enzymes typically occurs with the help of specialised units inside living cells. In most cases, the enzyme is unable to fold back into its original three-dimensional structure after becoming destabilised. Instead, the enzyme refolds into a more stable, but incorrect, structure and loses its ability to catalyse chemical reactions. Denaturation is therefore irreversible as the specific three-dimensional shape of the enzyme cannot be recovered.

Study tip

Changes to pH affect enzyme activity by modifying the types of interactions that side groups can form. If a functional group gains or loses a charge, this can affect its ability to form ionic bonds. This, in turn, changes the enzyme's three-dimensional shape.

Effect of pH on enzyme activity

Most enzymes in the human body work best at around a neutral pH of 7.

pH changes can cause enzymes to become denatured

As you learnt earlier, tertiary structures of enzymes can be held together by ionic interactions between charged side groups. For these interactions to occur, the side groups must gain a charge by either accepting or donating a proton. This exchange of protons can only occur at a particular or optimal pH in the environment.

For example, polar acidic side groups such as aspartic acid can interact with polar basic side groups such as lysine to form ionic attractions. However, at non-optimal pH, these side groups may start to react with the environment instead of with each other. In a more acidic environment (lower pH), basic side groups such as lysine react with H^+ ions in the environment. They gain a proton. In more alkaline environments (higher pH), acidic side groups such as aspartic acid react with OH^- ions in the environment. They lose a proton.

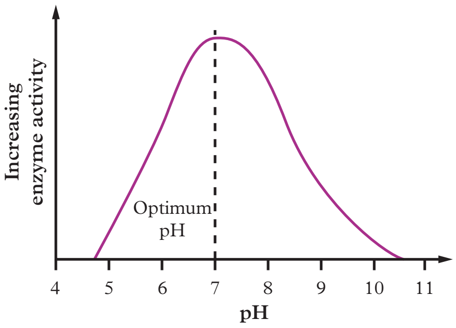


FIGURE 7 Effect of pH on enzyme activity

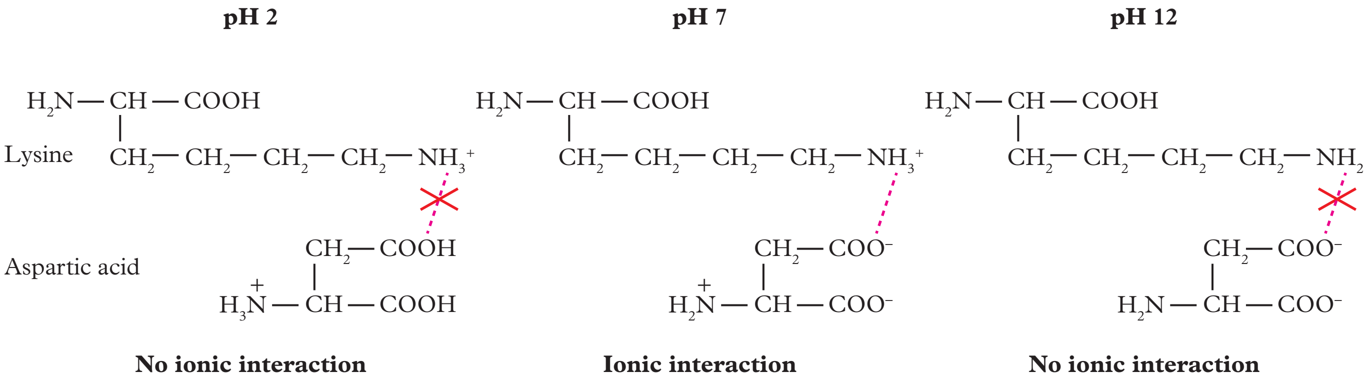


FIGURE 8 Lysine and aspartic acid only form ionic interactions around pH 7 because both NH_3 and COOH groups in the side chain become ionised.

If these groups are involved in ionic interactions to form the tertiary structure, they may become disrupted if they lose their charge due to a change in pH. In the same way, if they are usually uncharged in the original enzyme and gain a charge in a new pH environment, they may form new ionic interactions. This may alter the three-dimensional structure of the enzyme and cause it to become denatured. It may also affect how the enzyme binds to its substrate(s).

pH changes can cause enzymes to form zwitterions

As well as the side groups, the opposite ends of the polypeptide chain can become ionised. In a low pH environment (usually < 4), the $-NH_2$ group is likely to accept a proton and form a positively charged $-NH_3^+$. In a high pH environment (usually > 10), the $-COOH$ group is likely to lose a proton and form a negatively charged $-COO^-$ group.

At a pH range between these values, amino acids can form **zwitterions**. These molecules have a positive charge at one location on the molecule, and a negative charge at another location. However, these charges balance out so that the zwitterion has a net charge of zero. The acidic, zwitterionic and basic forms of an amino acid are shown in Figure 9. Worked example 12.4 shows you how to draw the three forms of methionine.

zwitterion
a molecule with a positively charged group and a negatively charged group

12.4 Worked example
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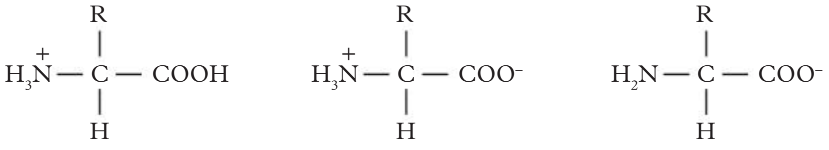
12.4 Worked example
Video demonstration



$$H_3\overset{+}{N}-\overset{\overset{\displaystyle R}{|}}{\underset{\underset{\displaystyle H}{|}}{C}}-COOH \qquad H_3\overset{+}{N}-\overset{\overset{\displaystyle R}{|}}{\underset{\underset{\displaystyle H}{|}}{C}}-COO^- \qquad H_2N-\overset{\overset{\displaystyle R}{|}}{\underset{\underset{\displaystyle H}{|}}{C}}-COO^-$$

Acid form (low pH) **Zwitterion (pH ~ 7)** **Base form (high pH)**

FIGURE 9 Acidic, zwitterionic and basic forms of an amino acid

12.4 CHECK YOUR LEARNING

Describe and explain

1 Explain what an enzyme is.

2 Complete the table below to summarise the four types of enzyme structure.

Structure	Description
Primary structure	
Secondary structure	
Tertiary structure	
Quaternary structure	

Apply, analyse and compare

3 Identify two amino acids with side chains that could form hydrogen bonds with each other.

4 Identify two amino acids with side chains that could form ionic interactions with each other.

5 Draw the structural formula of leucine at:
   - a pH 1.0
   - b pH 7.0
   - c pH 13.0.

6 Explain how changes in pH can reduce an enzyme’s ability to catalyse biological reactions.

Design and discuss

7 Explain why enzymes work less effectively when:
   - a the temperature is very low
   - b the temperature is above the optimal temperature.

8 Suggest a reason why not all enzymes have a quaternary structure.

12.5 Medicines that function as competitive enzyme inhibitors

KEY IDEAS

In this topic, you will learn that:

- some medicines function as competitive enzyme inhibitors
- enzyme inhibition can be reversible or irreversible.

Disease often results from problems with chemical reactions occurring in the body. Many of these reactions are catalysed by enzymes. This makes enzymes attractive targets for medicines. In this topic, you will explore a specific class of medicines that work by inhibiting (stopping or restricting) enzyme activity.

Competitive enzyme inhibitors

Some medicines bind directly to the active sites of enzymes and decrease their catalytic activity. These are called **competitive enzyme inhibitors** because they compete with the substrate for binding to the active site of the enzyme. Because they occupy the active site and do not undergo a reaction, this prevents the natural substrate from binding and catalysis does not occur. Examples of medicines that act as competitive inhibitors include erythromycin, sildenafil, and atorvastatin.

competitive enzyme inhibitor
a compound that binds to the active site of an enzyme and prevents the binding of the natural substrate

12.5 Real-world chemistry
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Remember that for an enzyme and substrate to bind successfully, their shapes must be complementary. The same applies to competitive inhibitors. They often have a similar structure to the enzyme's target substrate. Competitive inhibitors will therefore bind using the same lock-and-key mechanism. Some of these competitive inhibitors bind reversibly, while others bind irreversibly.

Reversible inhibition

reversible inhibitor
an inhibitor that does not bind permanently to the enzyme

Reversible inhibitors bind to an enzyme's active site with hydrogen bonding, ionic interactions, and dispersion forces. These intermolecular forces in the enzyme–substrate complex can be broken relatively easily. This means that after inhibiting the enzyme for some time, the medicine can be released from the active site of the enzyme and the enzyme is free to catalyse reactions again.

An example is shown in Figure 2, where succinate is the natural substrate for an enzyme called succinate dehydrogenase. Binding of succinate to succinate dehydrogenase produces a new molecule called fumarate. An inhibitor, malonate, has a similar structure to succinate. It competes with succinate to bind to the same active site, but does not undergo a reaction. This makes it a reversible competitive inhibitor, as the enzyme can be released to bind to another substrate.

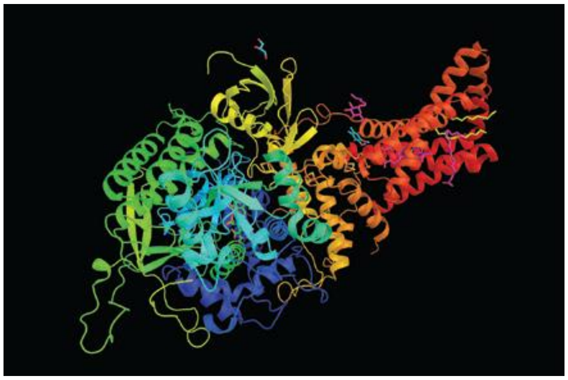


FIGURE 1 Succinate dehydrogenase has a complex three-dimensional structure.

12.5
PRACTICAL: LITERATURE REVIEW

Does the lock-and-key model fit?
Go to your **Student obook pro**.

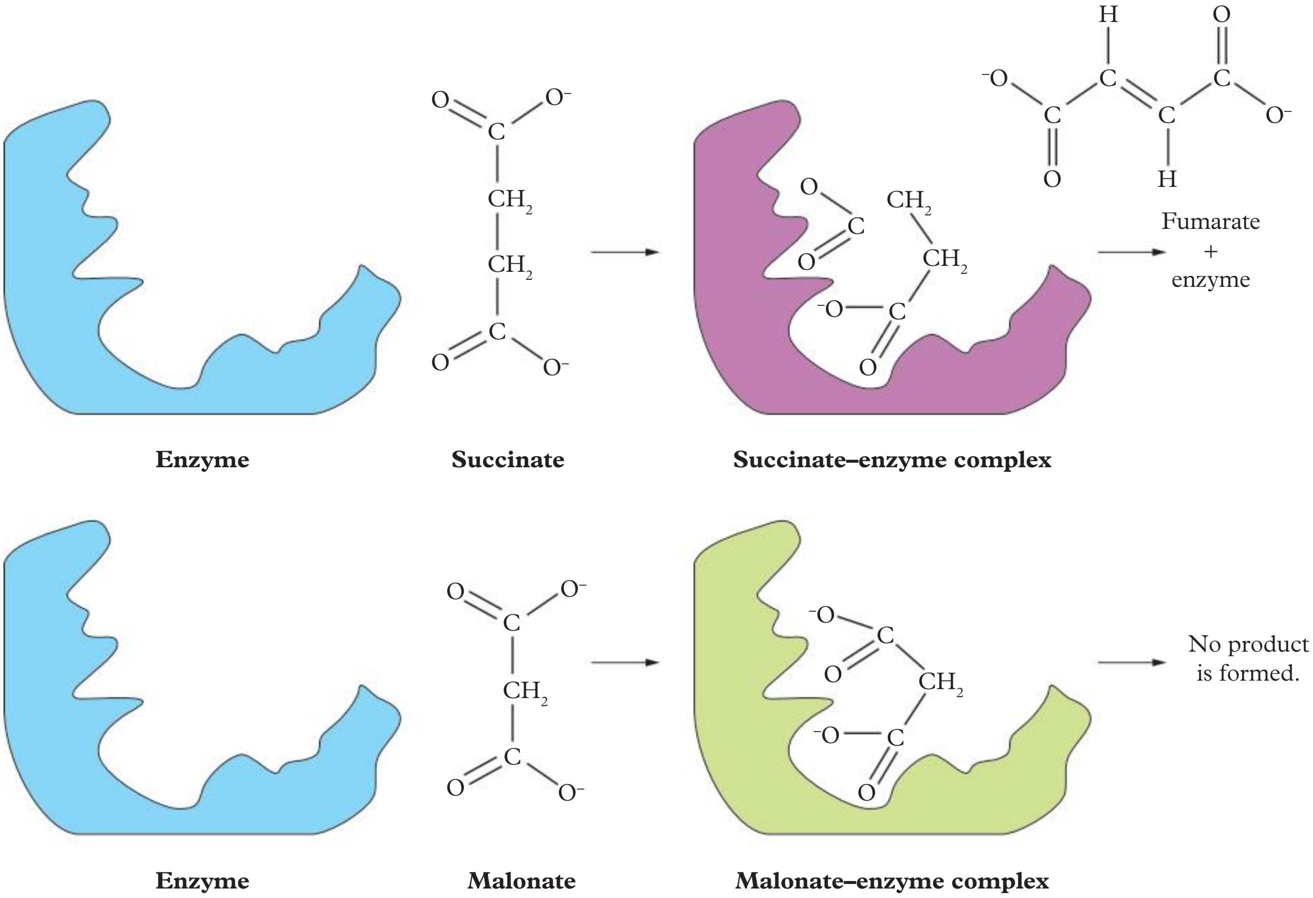


FIGURE 2 Malonate is a reversible competitive inhibitor of succinate dehydrogenase.

Irreversible inhibition

Irreversible inhibitors form stronger, usually covalent, bonds with the active site of the enzyme. Because the bonds cannot be easily broken, the active site of the enzyme is blocked permanently. Some irreversible inhibitors alter the structure of the enzyme (particularly at the active site) by adding or removing atoms from it. This causes permanent changes to the enzyme's three-dimensional shape and prevents it from binding with natural substrate molecules.

irreversible inhibitor
an inhibitor that binds permanently to the enzyme

An example is salicylic acid (the active ingredient in aspirin), which binds irreversibly to the cyclooxygenase 2 (COX-2) enzyme. COX-2 has a role in inflammation and pain. Salicylic acid adds an acetyl (CH_3CO) functional group to COX-2, preventing it from binding its natural substrates. Pretty handy if you have a bad headache!

12.5 CHECK YOUR LEARNING

Describe and explain

1 Explain how competitive enzyme inhibitors can be used as medicines.

2 Describe the characteristics of a competitive enzyme inhibitor.

3 Describe how the lock-and-key model applies to competitive enzyme inhibition.

Apply, analyse and compare

4 Compare reversible and irreversible enzyme inhibitors.

CHAPTER 12 Review

Chapter summary

12.1
- Bioactive compounds found in plants can be useful as medicines.
- Medicinal compounds can be extracted from plants using solvent extraction or distillation.

12.2
- The structure of an organic compound (its functional groups and how they are arranged) is important for its medicinal properties.

12.3
- The specific three-dimensional configuration of an organic compound can allow it to bind to targets in the human body.
- Organic compounds can exist as stereoisomers, which contain atoms connected in the same way but have a different three-dimensional configuration. This affects how they interact with biological targets.
- Stereoisomers contain at least one chiral carbon; i.e., a carbon with single bonds to four different atoms or groups of atoms.

12.4
- Enzymes are protein-based catalysts that speed up chemical reactions in the body.
- In catalysis, substrates bind to the active site of enzymes. This can be described using the lock-and-key model, where the shapes of the substrate and enzyme active site are complementary to each other.
- The specific three-dimensional shape of an enzyme's active site is determined by its primary, secondary, tertiary and quaternary structure.
- At low temperatures, enzyme activity is decreased because the number of successful collisions between the enzyme active site and substrate is decreased. At high temperatures, enzymes become denatured because heat disrupts the bonds holding together the enzyme's three-dimensional shape.
- In a pH environment outside of the optimal range, enzyme activity is reduced because the bonds holding the enzyme's three-dimensional shape may be disrupted.

12.5
- Some medicines are competitive enzyme inhibitors that block the active site of an enzyme and prevent its natural substrate from binding.
- Enzyme inhibition by competitive inhibitors can be reversible or irreversible.

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can...	Confidently	Mostly	Not really	Revision link
• explain how natural plant compounds that are possible active ingredients for medicines are extracted and purified, using solvent extraction and distillation	☐	☐	☐	Go back to Topic 12.1
• identify the structure and functional groups of organic molecules that are medicines	☐	☐	☐	Go back to Topic 12.2
• identify chiral centres in an organic molecule	☐	☐	☐	Go back to Topic 12.3
• explain the significance of isomers and chiral centres in the effectiveness of medicines	☐	☐	☐	Go back to Topic 12.3
• explain the function of enzymes as protein-based catalysts, including the lock-and-key model	☐	☐	☐	Go back to Topic 12.4
• describe the four levels of organisation of enzymes, including primary, secondary, tertiary and quaternary structure	☐	☐	☐	Go back to Topic 12.4
• explain how increased temperature, decreased temperature and changes to pH affect enzyme function	☐	☐	☐	Go back to Topic 12.4
• describe how medicines can function as competitive enzyme inhibitors	☐	☐	☐	Go back to Topic 12.5

Revision questions

Multiple choice

1 Which of the following compounds is a natural pesticide found in plants?

A Catechins
B Caffeine
C Nicotine
D All of the above

2 Which of these is a non-polar solvent?

A Hexane
B Butan-1-ol
C Water
D Propanone

3 Which statement about distillation is correct?

A Distillation separates compounds in a mixture by their molecular mass.
B Distillation separates compounds in a mixture by their boiling point.
C Distillation decreases the molecular mass of compounds in a mixture.
D Distillation decreases the purity of compounds in a mixture.

4 A chiral carbon atom is:

A a carbon atom bonded to four other carbon atoms.
B a carbon atom with no C–H covalent bonds.
C a carbon atom in a carbon–carbon double bond (C=C).
D a carbon atom bonded to four different atoms or groups of atoms.

5 Carvone contains a chiral carbon atom. The two forms of carvone, (R)-carvone (which smells like mint) and (S)-carvone (which smells like caraway) are:

A stereoisomers.

B cis/trans isomers.

C structural isomers.

D heavy isotopes.

FIGURE 1 Caraway plant

6 Enzymes are:

A solid substances that increase the activation energy for a biological reaction.

B biological substances that decrease the rate of a biological reaction.

C molecules that catalyse a biological reaction.

D liquids that increase the ΔH for a biological reaction.

7 The primary structure of a protein contains:

A ionic interactions between amino groups and carboxyl groups of amino acids.

B dispersion forces between amino acids.

C hydrogen bonds between polar amino acid side groups.

D covalent bonds between amino acids.

8 The secondary structure of a protein is:

A hydrogen bonds between adjacent amino acid side groups.

B hydrogen bonding between C=O of an amide group and N–H of a nearby amide group.

C hydrogen bonding between sulfur-containing amino acid side groups.

D hydrogen bonding between amine groups and water.

9 When the temperature is increased beyond its optimal temperature, enzyme activity:

A increases because the substrate and enzyme collide more frequently.

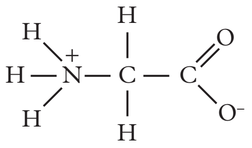
B decreases because the bonds in the enzyme–substrate complex are destabilised.

C decreases because the bonds holding together the enzyme's three-dimensional structure are destabilised.

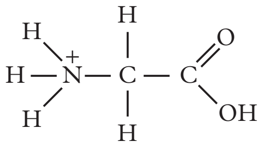
D decreases because the covalent bonds in an enzyme are broken.

10 Which of the following structures shows lysine at pH 2.0?

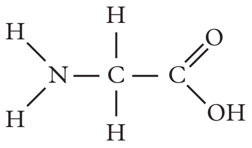
A



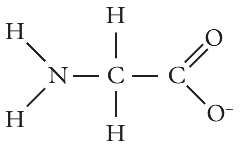
B



C



D



Short answer

Describe and explain

11 Define the term *zwitterion*.

12 Describe the role of nicotine as a secondary metabolite in tobacco plants.

13 Identify the properties of natural plant compounds that can be used to extract them from the plant.

14 Explain the importance of the three-dimensional structure in the effectiveness of medicines.

15 Explain why it is important to know if an organic compound to be used as a medicine has stereoisomers.

16 Enzymes work best at optimal conditions.

- **a** Explain why enzymes are less effective at low temperatures. Refer to collision theory in your answer.
- **b** Explain why enzymes are less effective when the temperature is greater than the optimal temperature. In your answer, refer to the types of chemical bonds broken during denaturation.
- **c** Explain why enzymes are less effective at non-optimal pH. In your answer, refer to the types of bonds in an enzyme's tertiary structure that are affected by changes in pH.

17 Aspartic acid is an amino acid. Draw the chemical structure of aspartic acid at pH 12.0.

FIGURE 2 Walnuts are rich in aspartic acid.

18 Explain why competitive enzyme inhibitors often have a similar structure to the natural substrate for the enzyme.

Apply, analyse and compare

19 Compare the overall charge of lysine at pH 2.0, 7.0 and 12.0.

20 Draw the zwitterions of the following amino acids.

- **a** Histidine
- **b** Tyrosine
- **c** Leucine
- **d** Threonine

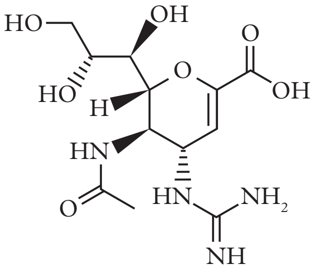
21 Draw the three-dimensional structural formulas of both stereoisomers of alanine. Indicate the chiral centre by writing an asterisk (*) next to it.

FIGURE 3 Poultry is an excellent source of alanine.

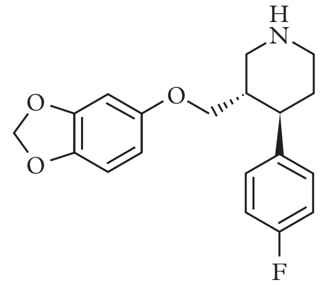
22 For the following medicinal compounds:

- **a** Identify the functional groups (halogens, carboxyl, hydroxyl, primary amine and amide groups only).
- **b** Identify the chiral centres by writing an asterisk (*) next to them.

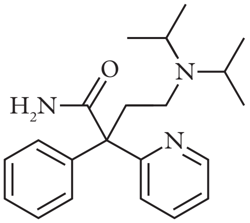
I Zanamivir



II Paroxetine



III Disopyramide



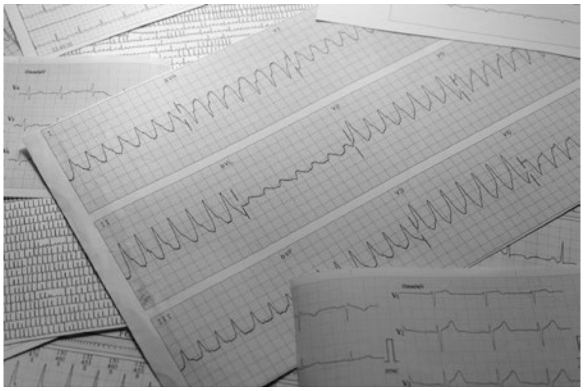
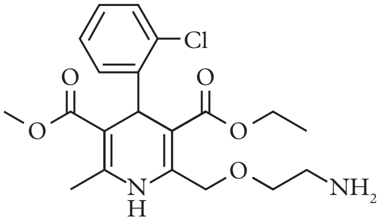


FIGURE 4 Disopyramide is used to treat arrhythmia (irregular heartbeat).

23 The molecular structure of amlodipine is shown below.



- **a** Circle and label two ester groups, a primary amine group and a secondary amine group on the structure of amlodipine.
- **b** Indicate any chiral carbon atoms by writing an asterisk (*) next to them.

24 Compare the primary, secondary and tertiary structures of enzymes.

25 Inhibitors decrease enzyme activity.

- **a** Explain how reversible inhibitors decrease enzyme activity.
- **b** Explain how irreversible inhibitors decrease enzyme activity.

Design and discuss

26 A chemist wants to extract and purify a new plant compound, which they believe can act as an antibiotic. They know that the compound is highly soluble in ethanol and has a boiling point of 230°C.

- **a** Design a method to separate the plant compound from the rest of the plant using solvent extraction.
- **b** Design a method to separate the plant compound from the rest of the plant using distillation.
- **c** Explain what other properties of the plant need to be known to successfully extract a pure sample of the plant compound.

27 A pharmacologist wants to design an experiment to test the effectiveness of a new medicine. The medicine is a competitive inhibitor for an enzyme called dihydrofolate reductase, which is involved in the replication of cancer cells in humans. She plans to incubate the medicine with the enzyme and observe the effect on enzyme activity.

- **a** Discuss the conditions that will be needed to carry out the experiment. Justify your response.
- **b** On the day of the experiment, the pharmacologist finds that the laboratory has run out of the solution of the required pH. Another pharmacologist on her team suggests using an alternative solution, with a pH 4 value greater than what she needs. The second pharmacologist assures her that it will not affect her experiment. Evaluate this claim and decide whether the pharmacologist should run the experiment.

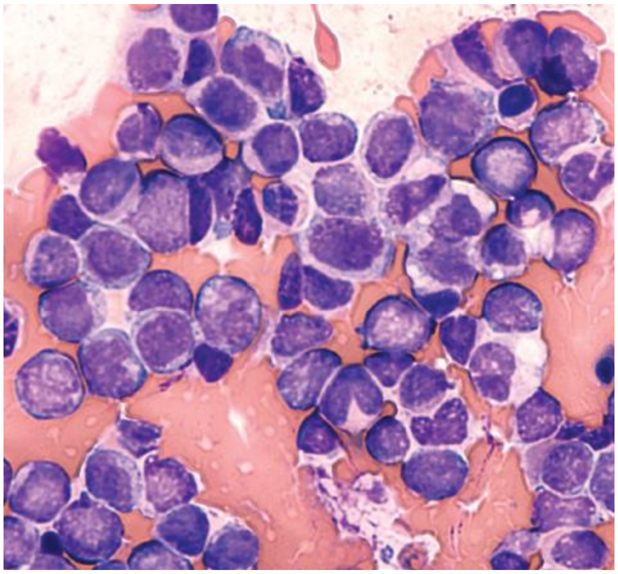


FIGURE 5 Cancer cells are important targets for modern medicines.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

AREA OF STUDY 2

UNIT 4 Checkpoint

Multiple choice

Question 1

Three isomers of the alcohol C_4H_9OH are tested using the purple-coloured potassium permanganate, to see if they will undergo oxidation. Which of the following solution colours would be observed when potassium permanganate is added?

	butan-2-ol	2-methyl propan-1-ol	2-methyl propan-2-ol
A	purple	colourless	colourless
B	colourless	colourless	purple
C	colourless	purple	purple
D	purple	purple	purple

Question 2

A 6.00 g sample of a hydrocarbon with a molar mass of 92 $g\,mol^{-1}$ is investigated to determine its degree of unsaturation. The sample if found to react with 33.09 g of I_2. The degree of unsaturation of the molecule is:

A 0

B 1

C 2

D 3

Question 3

A mixture containing five alcohols is separated using fractional distillation. The boiling points of the alcohols are listed in the table.

Propanol	Butanol	Pentanol	Hexanol	Heptanol
97°C	118°C	138°C	157°C	175°C

When the mixture is heated to 120°C, the fraction that remains in the distilling flask will contain:

A propanol and butanol.

B pentanol, hexanol and heptanol.

C butanol, pentanol, hexanol and heptanol.

D propanol, butanol, pentanol, hexanol and heptanol.

Question 4

A student performs a redox titration to determine the amount of ascorbic acid in an orange juice sample. They titrated 25.00 mL aliquots of the orange juice with a known concentration of potassium dichromate solution. An average titre of 19.70 mL of potassium dichromate was recorded after three titrations. Which of the following errors will produce an overestimation of the concentration of ascorbic acid?

A Using methyl red indicator

B Rinsing the conical flask with water

C Rinsing the pipette with water

D Rinsing the burette with water

Question 5

Aqueous ethanal is oxidised in aqueous acidic solutions of potassium dichromate, as shown by the equation:

$$3CH_3CHO(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow$$
$$3CH_3COOH(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$$

A 100.0 mL aliquot of ethanal required an average titre of 58.90 mL of a standard 0.0500 M aqueous potassium dichromate solution.

The concentration of the aldehyde in this solution is:

A 0.0885 M

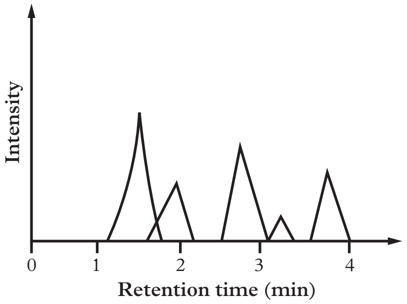
B 0.0442 M

C 0.0295 M

D 0.00982 M

Question 6

A sample containing five organic molecules is run through the HPLC. It uses a non-polar stationary phase and a polar mobile phase. The resulting chromatogram is shown below.



Which is the most likely order for the peaks, from shortest to longest retention time?

A Hexane, fluorohexane, hexanamine, hexanol, hexanoic acid

B Hexane, hexanol, hexanamine, fluorohexane, hexanoic acid

C Hexanoic acid, fluorohexane, hexanol, hexanamine, hexane

D Hexanoic acid, hexanol, hexanamine, fluorohexane, hexane

Question 7

Identify the molecule that produces ^{1}H-NMR spectrum with a different number of peaks to its ^{13}C-NMR spectrum.

A 2-methylpentan-2-ol

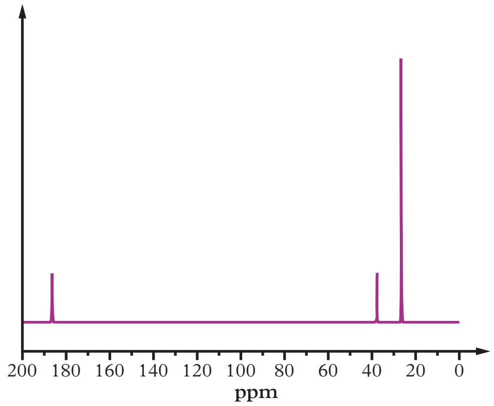
B 3,3-dimethylbutan-1-ol

C 4-methylpentan-1-ol

D 3-methylpentan-3-ol

Question 8

Which molecule does the following ^{13}C-NMR spectrum represent?



A $CH_3CH_2CH_3$

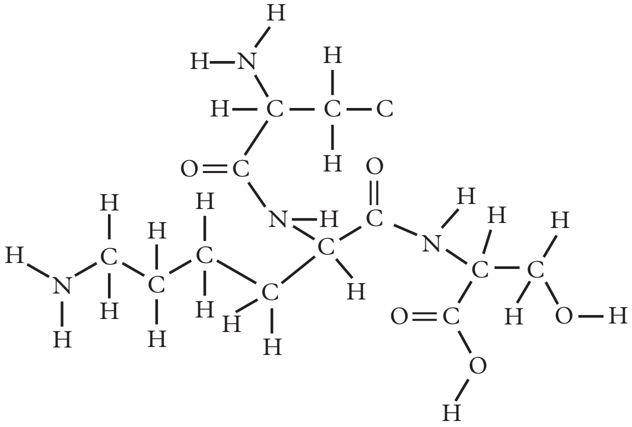
B $(CH_3)_3CCOOH$

C $(CH_3)_2CHCl$

D $CH_3CH_2CH_2OH$

Question 9

The structure below is a section of enzyme.



It is a:

A primary structure consisting of alanine, lysine and serine.

B secondary structure consisting of alanine, lysine and serine.

C secondary structure consisting of glutamine, leucine and aspartic acid.

D primary structure consisting of glutamine, leucine and aspartic acid.

Question 10

Which two amino acids hydrogen bond to form the tertiary structure of an enzyme?

A Serine and methionine

B Valine and lysine

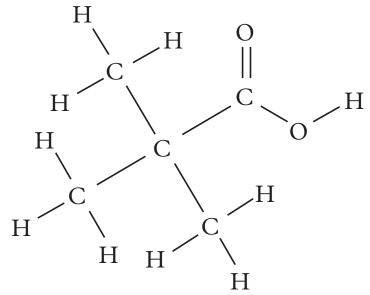
C Lysine and asparagine

D Cysteine and glutamic acid

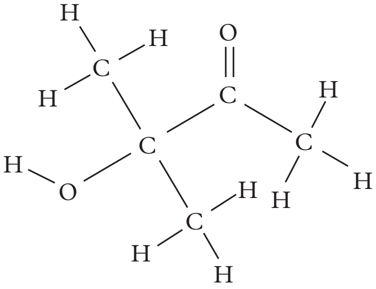
Short answer

Question 1 (13 marks)

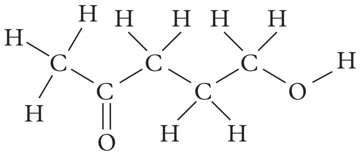
Four molecules are shown below.



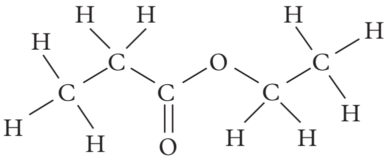
Molecule D



Molecule E

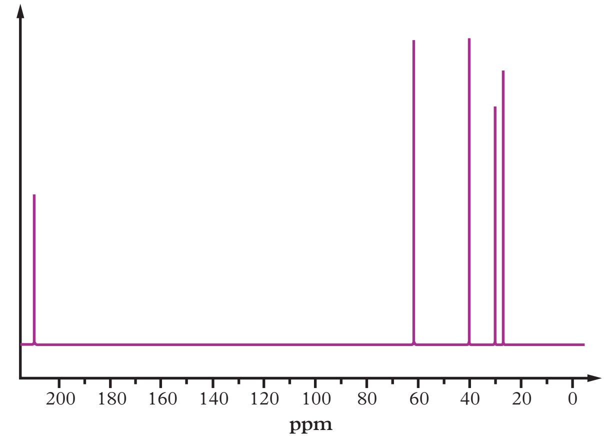


Molecule F



Molecule G

a A ^{13}C-NMR spectrum is shown below.

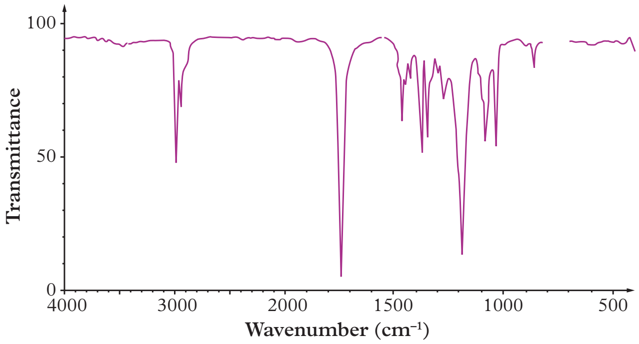


i Identify which of the four molecules produced this spectrum. 1 mark

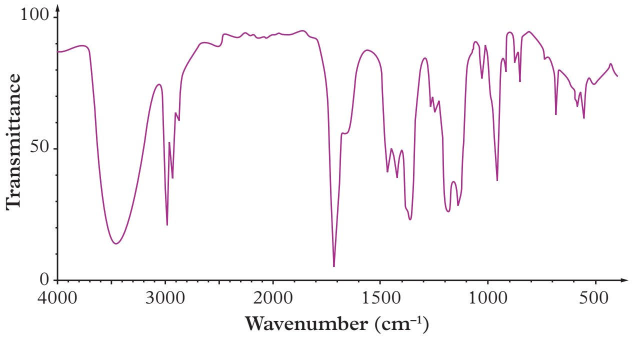
ii Explain your reasoning for selecting this molecule. 1 mark

b Two IR spectra, A and B, are shown below.

Spectrum A

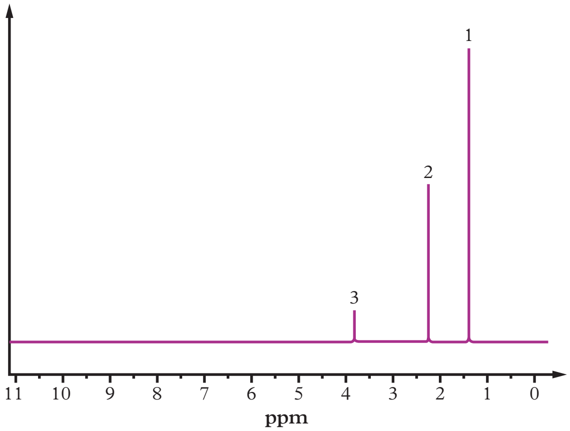


Spectrum B



- **i** Identify which of the four molecules produced spectrum A. 1 mark
- **ii** Explain your reasoning for selecting this molecule. 1 mark
- **iii** Discuss whether you can definitively determine which molecule produced spectrum B. 2 marks

c The high resolution ^{1}H-NMR for one of the molecules is shown below.

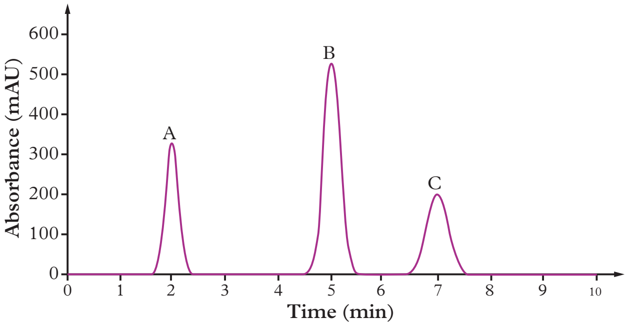


- **i** Identify which of the four molecules produced this spectrum. 1 mark
- **ii** Identify which peak belongs to which hydrogen environment in the molecule you have selected. 1 mark
- **iii** Describe the number of peaks you would expect to see for each of the other molecules. 3 marks

d Discuss whether you could use mass spectroscopy to determine the difference between the four molecules. 2 marks

Question 2 (6 marks)

To separate three performance-enhancing drugs, HPLC was completed using a polar stationary phase and a non-polar mobile phase. The following chromatogram was obtained. This can be used as a standard to compare samples of blood from athletes to check for drug use.

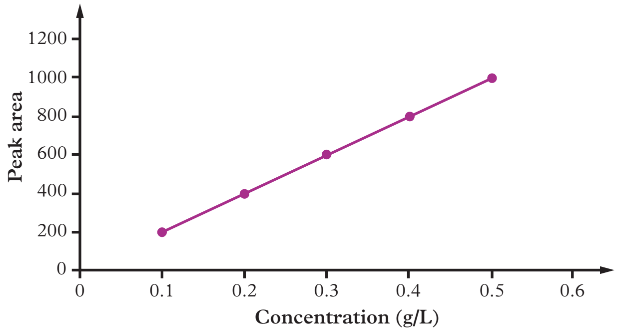


a Analyse the chromatogram and compare the molar masses of the three performance-enhancing drugs. 2 marks

b Discuss the polarity of the drug that produced the peak with the longest retention time. 1 mark

c A sample of an athlete's blood was run using a HPLC. The blood contained a peak at five minutes. Explain what this peak means for the blood sample. 1 mark

The following calibration curve was created using known concentrations of the same drug that produced the peak at five minutes.

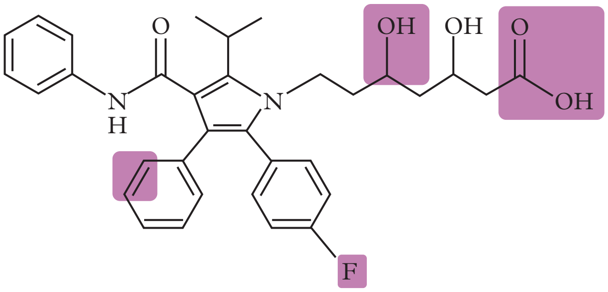


The peak area for the peak in the athlete's blood sample was 700.

d Calculate the concentration of the drug in the sample in mg/L. 2 marks

Question 3 (10 marks)

The structure of the drug atorvastatin, used as a heart medication, is shown.



a Name the four functional groups that are highlighted. 4 marks

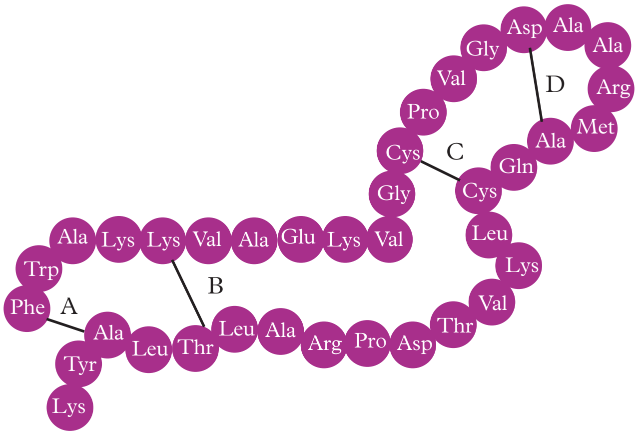
b **i** The drug atorvastatin has two chiral carbons. Identify them on the structure. 2 marks

ii Explain what a stereoisomer is. Identify the number of stereoisomers of atorvastatin. 2 marks

c Discuss how it is possible that only one of the stereoisomers is biologically active as a heart medication. 2 marks

Question 4 (11 marks)

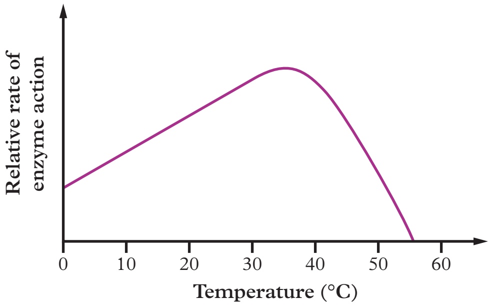
A section of an enzyme is shown below.



a Explain which structure is supported by the four bonds labelled A–D. 1 mark

b Identify the type of bonds labelled as A–D. 4 marks

The effect of temperature on the activity of the enzyme is shown in the graph on the right.



c From the graph, identify the optimal temperature for the enzyme. 1 mark

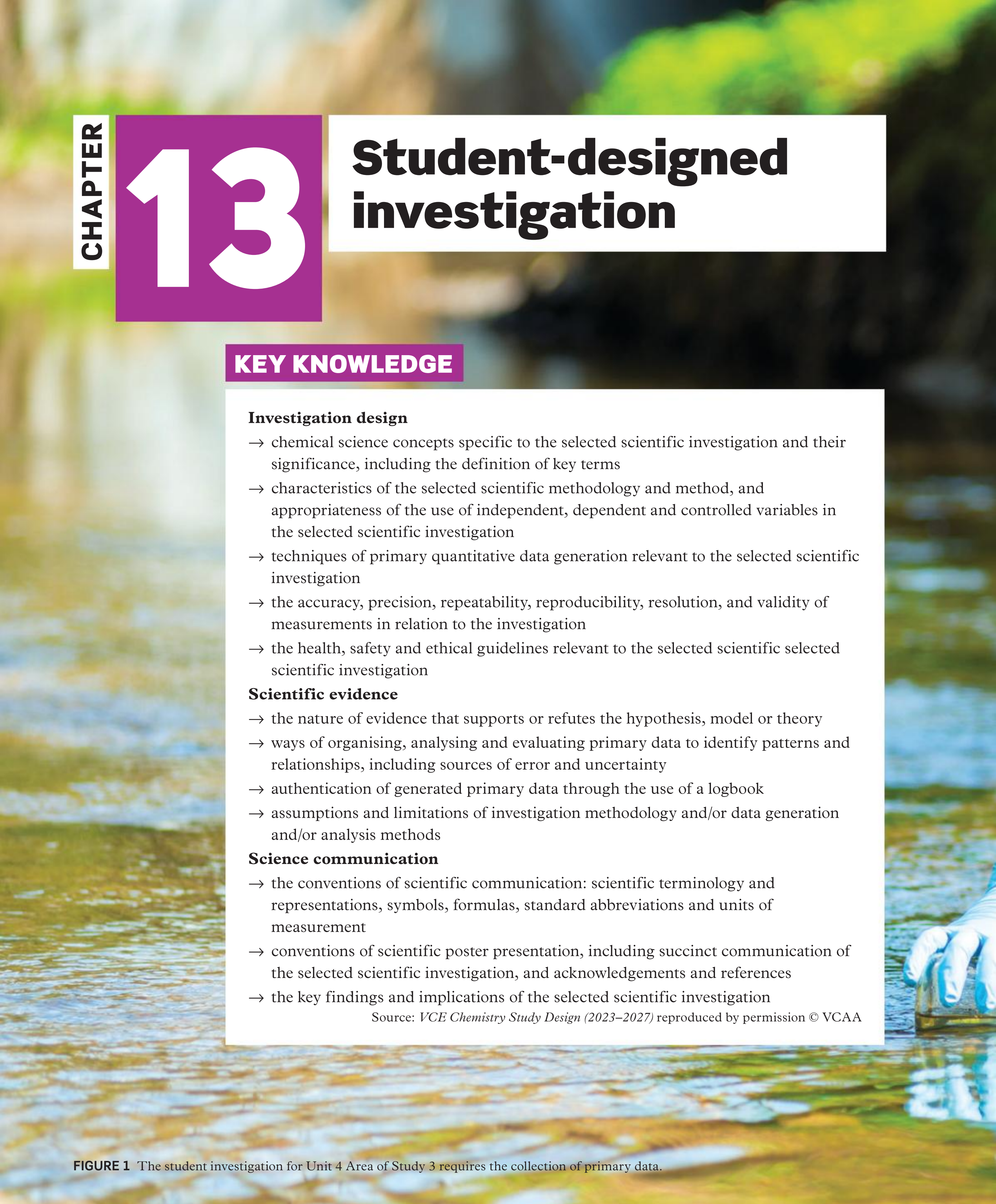
d Explain what would happen to the enzyme if it were used at a temperature of 10°C. 1 mark

e Explain how some medicines can act as competitive inhibitors and describe the properties they need to do this. 2 marks

f Compare the bonding of a reversible competitive inhibitor to that of an irreversible competitive inhibitor. 2 marks

TOTAL MARKS

/50 marks



CHAPTER 13

Student-designed investigation

KEY KNOWLEDGE

Investigation design

- → chemical science concepts specific to the selected scientific investigation and their significance, including the definition of key terms
- → characteristics of the selected scientific methodology and method, and appropriateness of the use of independent, dependent and controlled variables in the selected scientific investigation
- → techniques of primary quantitative data generation relevant to the selected scientific investigation
- → the accuracy, precision, repeatability, reproducibility, resolution, and validity of measurements in relation to the investigation
- → the health, safety and ethical guidelines relevant to the selected scientific selected scientific investigation

Scientific evidence

- → the nature of evidence that supports or refutes the hypothesis, model or theory
- → ways of organising, analysing and evaluating primary data to identify patterns and relationships, including sources of error and uncertainty
- → authentication of generated primary data through the use of a logbook
- → assumptions and limitations of investigation methodology and/or data generation and/or analysis methods

Science communication

- → the conventions of scientific communication: scientific terminology and representations, symbols, formulas, standard abbreviations and units of measurement
- → conventions of scientific poster presentation, including succinct communication of the selected scientific investigation, and acknowledgements and references
- → the key findings and implications of the selected scientific investigation

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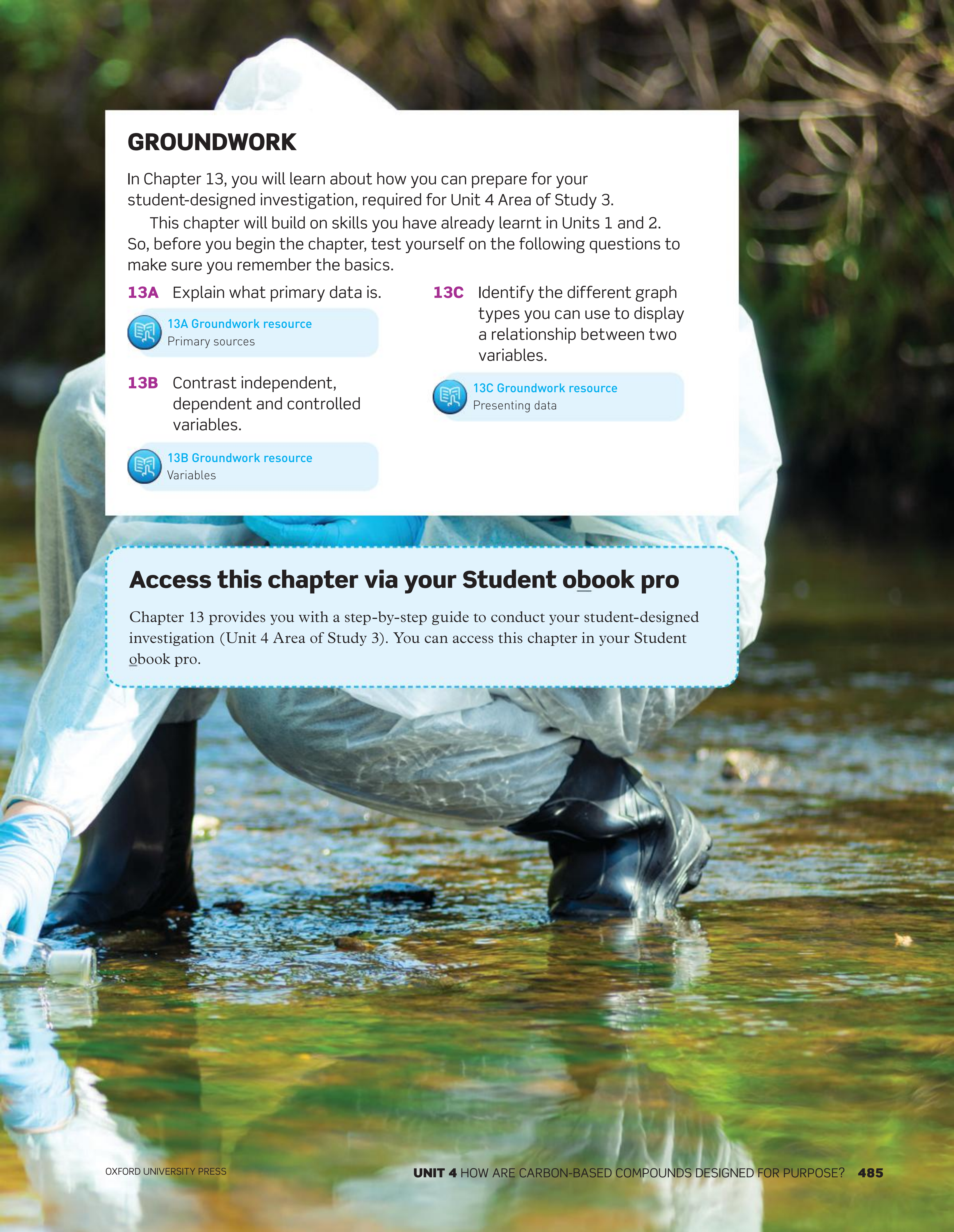
FIGURE 1 The student investigation for Unit 4 Area of Study 3 requires the collection of primary data.



GROUNDWORK

In Chapter 13, you will learn about how you can prepare for your student-designed investigation, required for Unit 4 Area of Study 3.

This chapter will build on skills you have already learnt in Units 1 and 2. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

13A Explain what primary data is.

13A Groundwork resource
Primary sources

13B Contrast independent, dependent and controlled variables.

13B Groundwork resource
Variables

13C Identify the different graph types you can use to display a relationship between two variables.

13C Groundwork resource
Presenting data

Access this chapter via your Student obook pro

Chapter 13 provides you with a step-by-step guide to conduct your student-designed investigation (Unit 4 Area of Study 3). You can access this chapter in your Student obook pro.

13.1

Topic selection

KEY IDEAS

In this topic, you will:

- write a scientific question
- select an experiment design or adapt one
- conduct research and define key theories
- evaluate any sources used in your research
- understand how to analyse the assessment criteria.

This chapter will guide you through your student-designed investigation. Each topic will focus on the essentials of investigation design, from planning and conducting the investigation, to communicating your results. Goals and to-do lists for each stage of the investigation are also included to help you.

Before you get started, make sure you have the details of this assessment task from your teacher. You should always check what the assessment criteria will be and note any tips or instructions your teacher might provide to help you succeed.

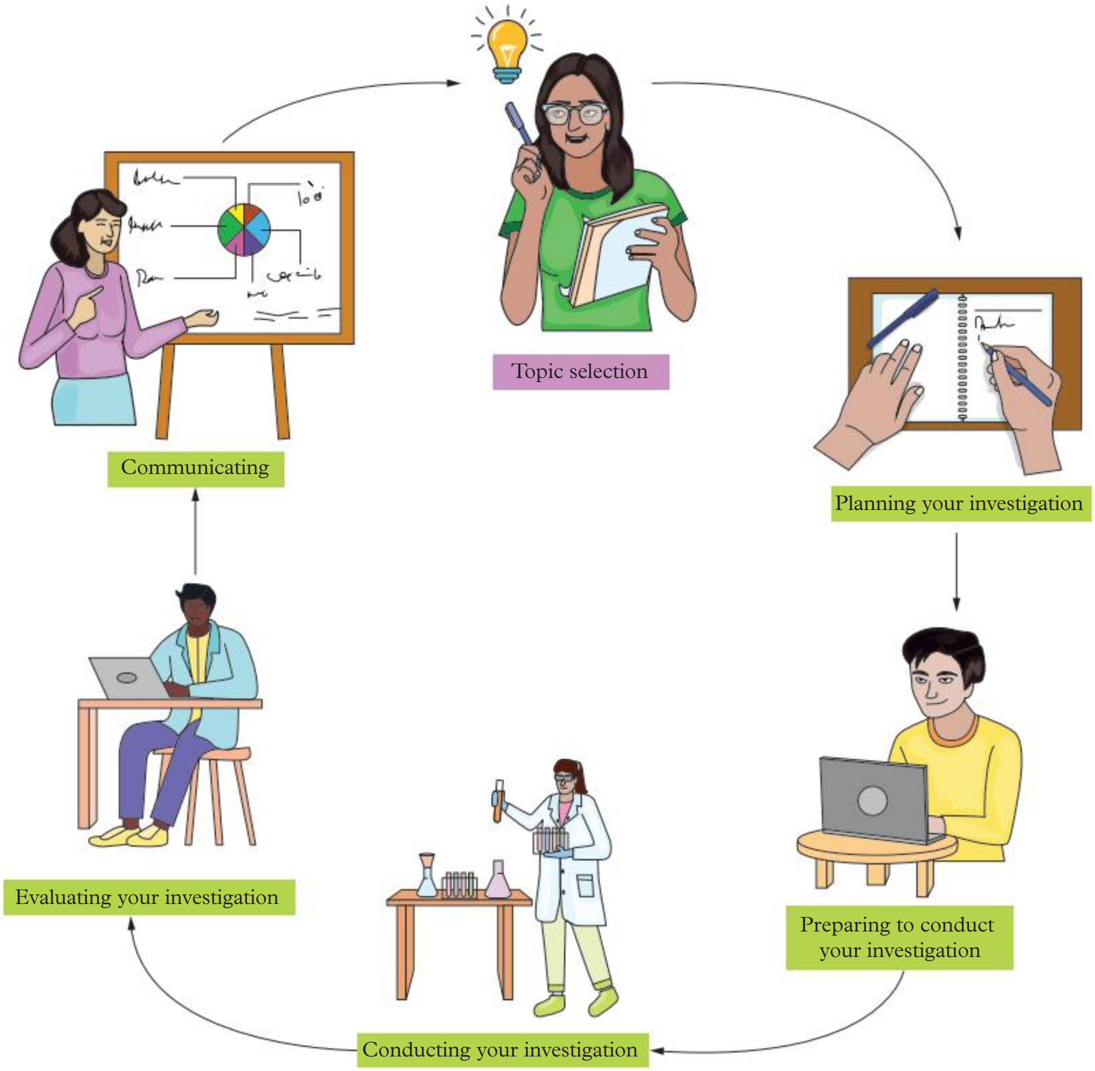


FIGURE 1 The scientific investigation process begins with the topic selection phase.

A scientific investigation is a multistep process. In Figure 1, you can see that the process can cycle, specifically when the final steps of your first investigation prompt a second investigation. Following the scientific process is essential to the success of the investigation.

This chapter will go through each phase of the process so that you can complete your own scientific investigation for Unit 4 Area of Study 3: 'How is scientific inquiry used to investigate the sustainable production of energy and/or materials?'

The start of your investigation begins with the topic selection phase. This phase includes:

- deciding on your topic
- deciding to design or adapt an experiment
- choosing and writing an investigation question
- conducting research and defining key theories
- analysing the assessment criteria.

Deciding on your topic

Your teacher may have already assigned you a specific topic or set of topics to select from. However, if you are deciding on your own topic, then you must select one that relates to the production of energy and/or chemicals and/or the analysis or synthesis of organic compounds.

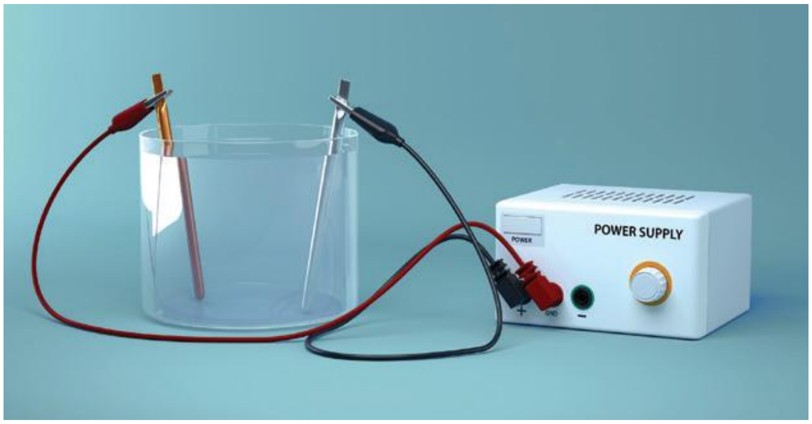


FIGURE 2 Your investigation topic must relate to concepts covered in Unit 3 and/or Unit 4.

Reading through Chapters 2–12 can help to refresh your memory of these topics. When deciding on your topic, you may also wish to consider investigating how chemistry concepts from one of these topics relates to the United Nations Sustainable Development Goals. Your investigation could relate to sustainability, or energy or material production.

FIGURE 3 Your research topic could integrate chemistry concepts and the sustainability perspectives, including the United Nations Sustainable Development Goals.

Topic selection can also be inspired by many other areas. Table 1 presents some potential sources for your topic selection and a thinking prompt you can use to help assess your potential interest in a topic.

TABLE 1 Sources and ideas to help inspire your topic selection

Potential source	Thinking prompt
Direct observation of and curiosity about an object, an event, a phenomenon, a practical problem or a technological development	Have any science concepts, real-world chemistry extracts or scientific articles gained your interest?
Anomalous or surprising investigation results or findings from analysis of qualitative and/or quantitative data	Throughout practicals that you have completed, has anything occurred that particularly spiked your interest or intrigue?
Extension of a previous inquiry	Have any inquiry tasks in practicals led you to ask further questions?
Research involving secondary data	Have you come across something that others have investigated that you might like to expand upon or look into further?

Whatever the source of inspiration for your topic, you should aim to select a topic that interests you. This will make the process more enjoyable and allow you to engage more with it. Topic selection can also occur while choosing and writing a question and/or designing or adapting an experiment.

Deciding to design or adapt an experiment

When you are looking at potential experiments to identify or test a research question, you should consider the following:

- Has your teacher outlined an experiment or set of experiments you have to use to complete your investigation?
- Did you complete an experiment during Units 3 and 4 that left you asking a question, or collect any data that did not quite align with your hypothesis?

If either of these is the case, then reusing or adapting an experiment that you have completed during Units 3 and 4 will be the best way to complete your investigation.

If you have a topic you intend to investigate and have not previously completed a practical activity on it, you can look online for experiments to use or adapt. You may also design your own experiment. Table 2 presents some of the different types of inquiry methods you can use to collect primary data.

If you are designing your own experiment, or reusing or adapting one, make sure to consider:

- whether you can complete the experiment in the time you have been given
- the equipment you have access to
- whether the experiment allows you to collect primary data.

TABLE 2 Inquiry methods that can be used to collect primary data

Inquiry type	Inquiry outline	Inquiry question types	Example questions
Controlled experiment	An investigation of the relationship between an independent variable (IV) and a dependent variable (DV), while controlling all other variables (CVs).	*What effect does … have on …?* *Is … related to …?*	What effect does the size of electrodes have on electroplating? Are the storage conditions of rechargeable batteries related to their efficiency?
Pattern seeking	The investigation of one variable to determine what other variable(s) can affect it, and to what extent other variables may be important in their effects on the variable under investigation. This can include observing natural events or phenomena and identifying patterns and/or relationships, and then proposing a link. It may involve multiple variables, because some variables may be difficult to control.	*What factors affect …?* *What are the optimal conditions for …?*	What factors affect the energy released from the combustion of fuels? What are the optimal conditions for a non-rechargeable cell to produce the highest voltage?
Single-variable exploration	The investigation of one variable or factor at a time, usually to see how it changes over time, focusing on observations and identification of a phenomenon. This type of inquiry can lead to questions about the causes of an observed phenomenon and prompt further investigation.	*How does … change over time?* *Do all …?* *When do(es) …?*	How does the voltage in a galvanic cell change over time? Do all plant oils create biodiesels with the same properties? When does the electrochemical series order change?

Choosing and writing an investigation question

A scientific inquiry or investigation involves asking or responding to an investigation question and then performing experiments and reporting on your findings. Your investigation and report should always respond to the question that you initially started with.

Once you have a topic or experiment in mind, you should start to consider questions that you have about that topic. You could also start with a general question that you have come across during Units 3 and 4 and narrow this down into a research question. How you refine your question can vary depending on whether you start off with a topic or idea or with an experiment in mind.

FIGURE 4 Selecting an investigation topic that interests you will make the investigation more worthwhile and enjoyable.

Starting with a topic or idea

Figure 5 shows a process that could help you to write an investigation question starting with a topic or idea. This process starts with the topic or idea, considers the theory involved and then structures questions around the relevant theory.

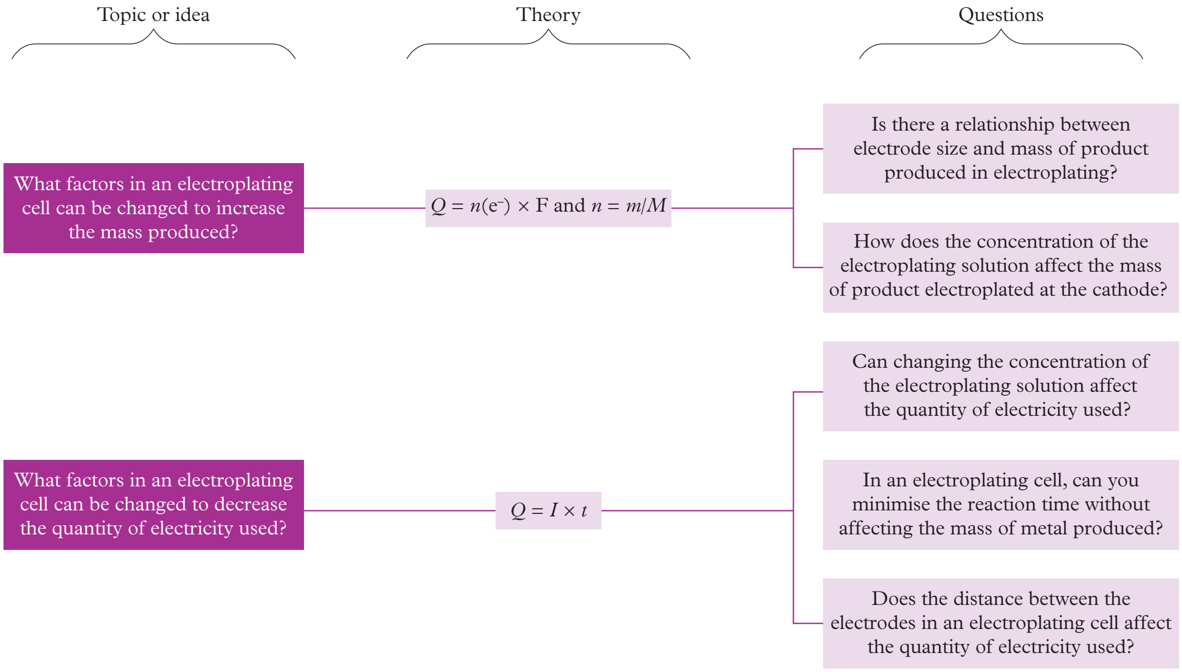


FIGURE 5 An example of the process involved in developing a research question from a topic or idea

Starting with an experiment

If your teacher has given you an experiment or a set of experiments to complete, or you completed an experiment and it left you with questions, you could use the process in Figure 6 to help build your question.

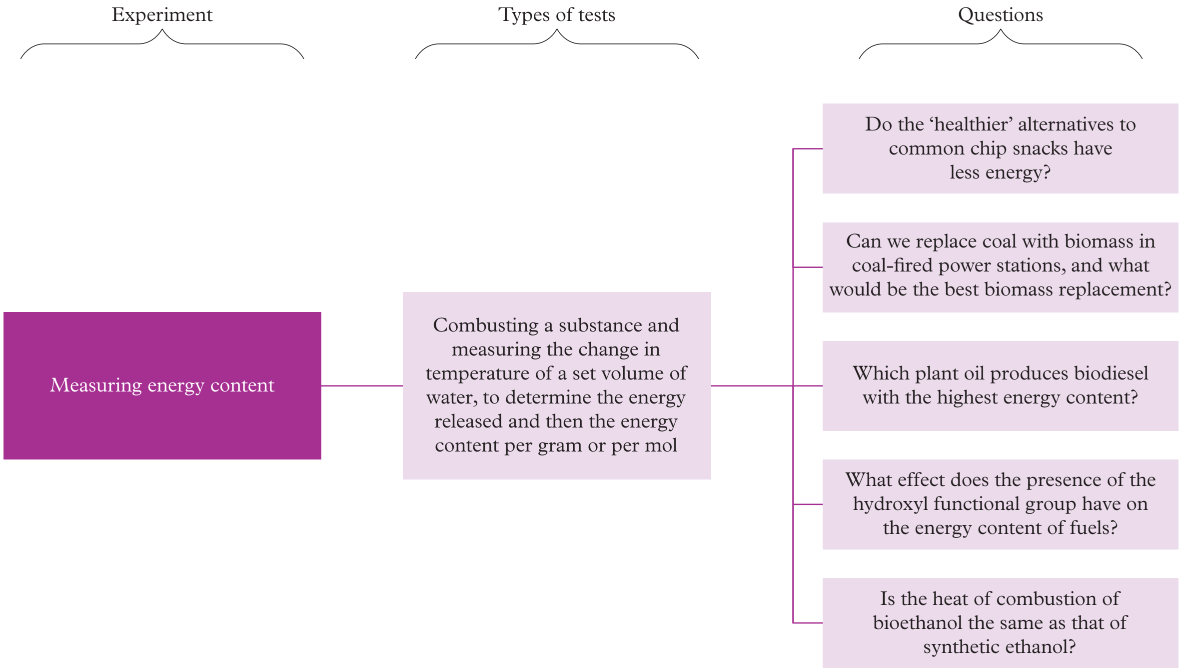


FIGURE 6 An example of the process involved in developing a research question from an experiment

After completing the processes in Figure 5 or 6, you might have a variety of questions to choose from. You might also need to go back a step and look at designing or adapting an experiment that will fit with your question and the time allocated. Whatever investigation question you decide on, it should:

- be clear and focused
- have an appropriate scope (not too vague or too narrow)
- not be too easy or difficult to answer (the question should require more than a simple yes or no answer)
- be researchable
- be analytical rather than descriptive.

Research tip

Always write your investigation question as a question and not as a statement.

Research tip

Read your proposed question to a friend or teacher and ask them if they think it delivers on all five of the key features of an investigation question.

Conducting research and defining key theories

Once you have your topic or experiment and you have a good scientific question, you should conduct some research into the theories surrounding your topic.

Research will help you to:

- understand the key theories or phenomena surrounding your topic
- find information to write your introduction
- determine better experiments or methods for conducting your experiment
- write your hypothesis and understand the reasoning behind your prediction
- better understand the data that you might need to collect
- fill any current gaps in your knowledge.

Before starting to research and compiling your key theories, read the information on sourcing information and analysing your chosen secondary sources by using the CRAAP method:

Currency: When was the information published? Is it out of date?

Relevance: How does the information fit in with your research question?

Authority: Who published or wrote this information? Are they qualified?

Accuracy: Is the information valid, reproducible, repeatable and accurate?

Purpose: Why does this information exist? Does the author have an agent or bias?

If a source does not pass the CRAAP test – for example, if it is out of date, or written by a biased source – do not use it to research.

Once you have confirmed that your sources are reliable, you can organise your information. There are many ways to do this. One example is to use dot points to summarise the information. For each dot point, you could make a note of how the information is relevant to your investigation question. You might also like to try a mind map, a table, Venn diagrams, or other graphic organisers.

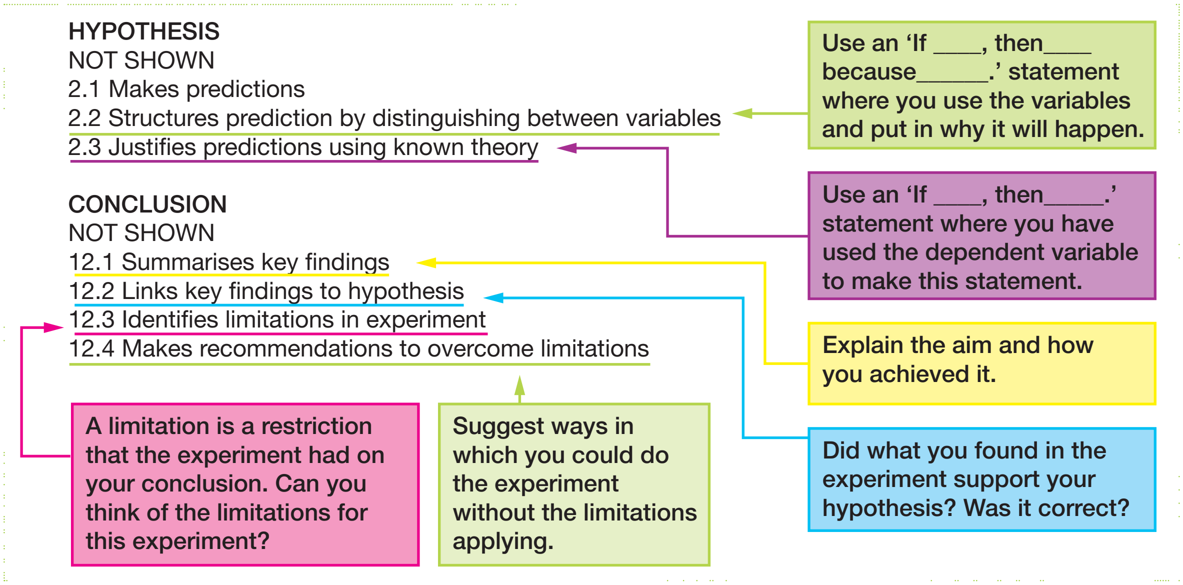
Drawing links between multiple sources also helps you to consolidate your understanding. It also acts as a way to critically examine the information you have collected. Some important questions to ask yourself are:

- Do the sources present the same or similar ideas?
- Is there conflicting information?
- Has any information been extended or built from previously established information?

FIGURE 7 Scientific magazines are a useful source of information for your research, but you should still evaluate them before you include any information from them.

Analysing the assessment criteria

When given your assessment criteria for your Unit 4 Area of Study 3 Scientific investigation, you should check to see what is required to achieve the highest marks. You can do this by annotating the rubric or criteria (Figures 8 and 9). You can also use the assessment criteria as a checklist (Figure 10).



HYPOTHESIS
NOT SHOWN
2.1 Makes predictions
2.2 Structures prediction by distinguishing between variables
2.3 Justifies predictions using known theory

CONCLUSION
NOT SHOWN
12.1 Summarises key findings
12.2 Links key findings to hypothesis
12.3 Identifies limitations in experiment
12.4 Makes recommendations to overcome limitations

Use an ‘If ____, then____ because______.’ statement where you use the variables and put in why it will happen.

Use an ‘If ____, then_____.’ statement where you have used the dependent variable to make this statement.

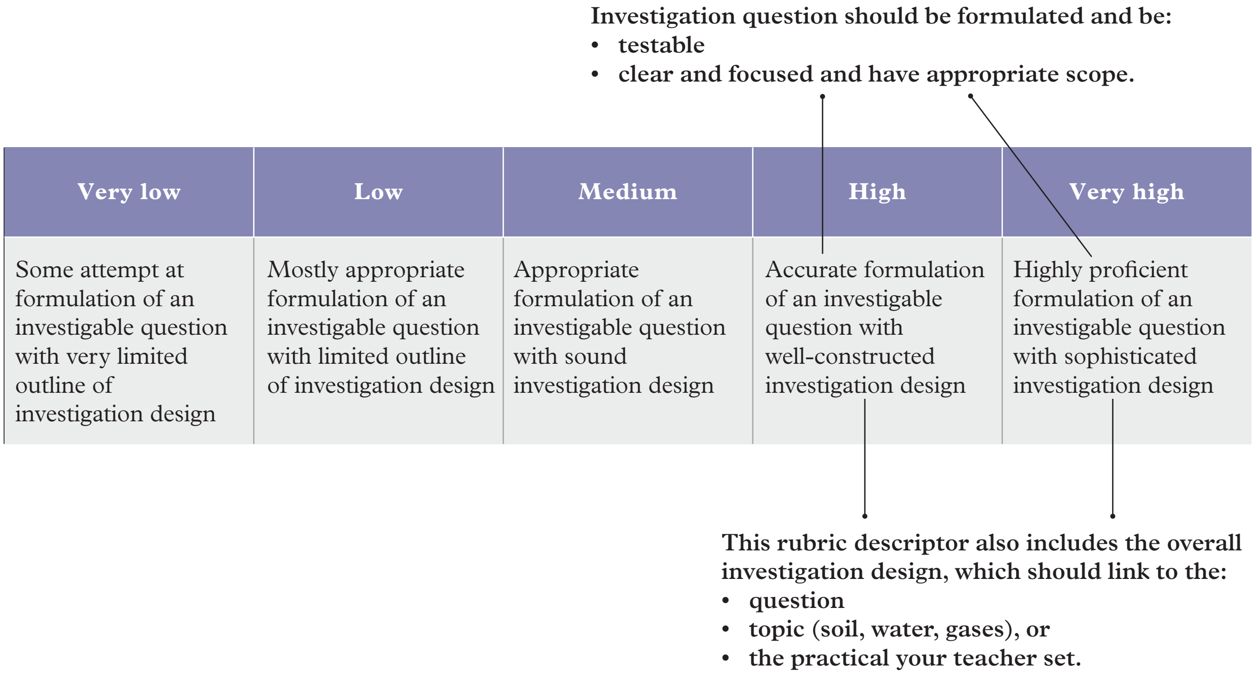
Explain the aim and how you achieved it.

Did what you found in the experiment support your hypothesis? Was it correct?

A limitation is a restriction that the experiment had on your conclusion. Can you think of the limitations for this experiment?

Suggest ways in which you could do the experiment without the limitations applying.

FIGURE 8 An example of a developmental rubric for a hypothesis and a conclusion, and annotations of what to do to get each mark

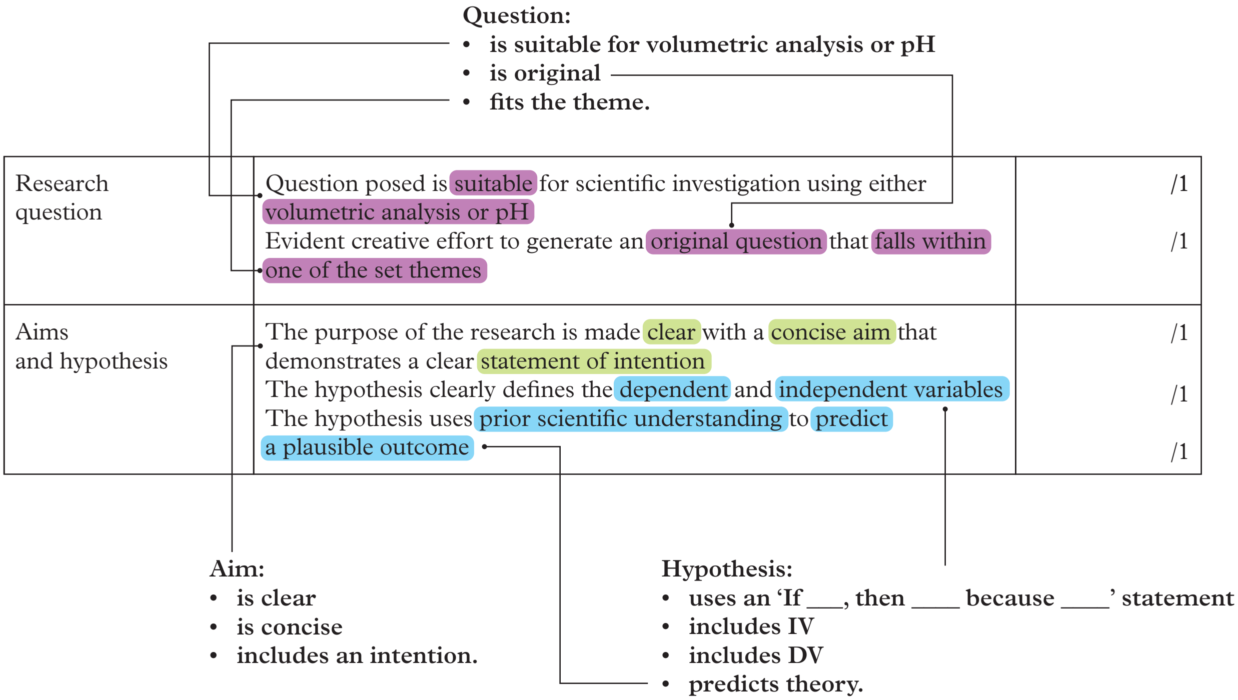


Investigation question should be formulated and be:
- **testable**
- **clear and focused and have appropriate scope.**

Very low	Low	Medium	High	Very high
Some attempt at formulation of an investigable question with very limited outline of investigation design	Mostly appropriate formulation of an investigable question with limited outline of investigation design	Appropriate formulation of an investigable question with sound investigation design	Accurate formulation of an investigable question with well-constructed investigation design	Highly proficient formulation of an investigable question with sophisticated investigation design

This rubric descriptor also includes the overall investigation design, which should link to the:
- **question**
- **topic (soil, water, gases), or**
- **the practical your teacher set.**

FIGURE 9 An example of a performance descriptor rubric for the investigation question and investigation design with annotations of the breakdown of the top marks



Question:
- **is suitable for volumetric analysis or pH**
- **is original**
- **fits the theme.**

Research question	Question posed is suitable for scientific investigation using either volumetric analysis or pH Evident creative effort to generate an original question that falls within one of the set themes	/1 /1
Aims and hypothesis	The purpose of the research is made clear with a concise aim that demonstrates a clear statement of intention The hypothesis clearly defines the dependent and independent variables The hypothesis uses prior scientific understanding to predict a plausible outcome	/1 /1 /1

Aim:
- **is clear**
- **is concise**
- **includes an intention.**

Hypothesis:
- **uses an 'If ___, then ____ because ____' statement**
- **includes IV**
- **includes DV**
- **predicts theory.**

FIGURE 10 An example of an assessment criteria rubric and annotations with checklists for each step

Remember to check through the assessment criteria you are given at the start to help you determine what you must do and include. You should also go through the criteria again at the end and double check that you have included everything.

FIGURE 11 Make sure you check the assessment criteria at the start and at the end of your investigation.

TO-DO LIST

- Write a scientific question that you would like to investigate.
- Make sure your question is posed as a question and not a statement.
- Select an experiment to adapt or design for your investigation.
- Define the key scientific theories and terms that are relevant to your investigation topic.
- Evaluate any sources used in your background research.
- Analyse the criteria you will be marked on for this assessment.

13.2

Planning your investigation

KEY GOALS

In this topic, you will:

- define your variables
- write a suitable hypothesis for your experiment
- construct a methodology for your experiment.

After you have selected your topic, the next phase involves planning your scientific investigation. This will include:

- defining the variables
- writing a hypothesis
- designing a method (including materials and equipment).

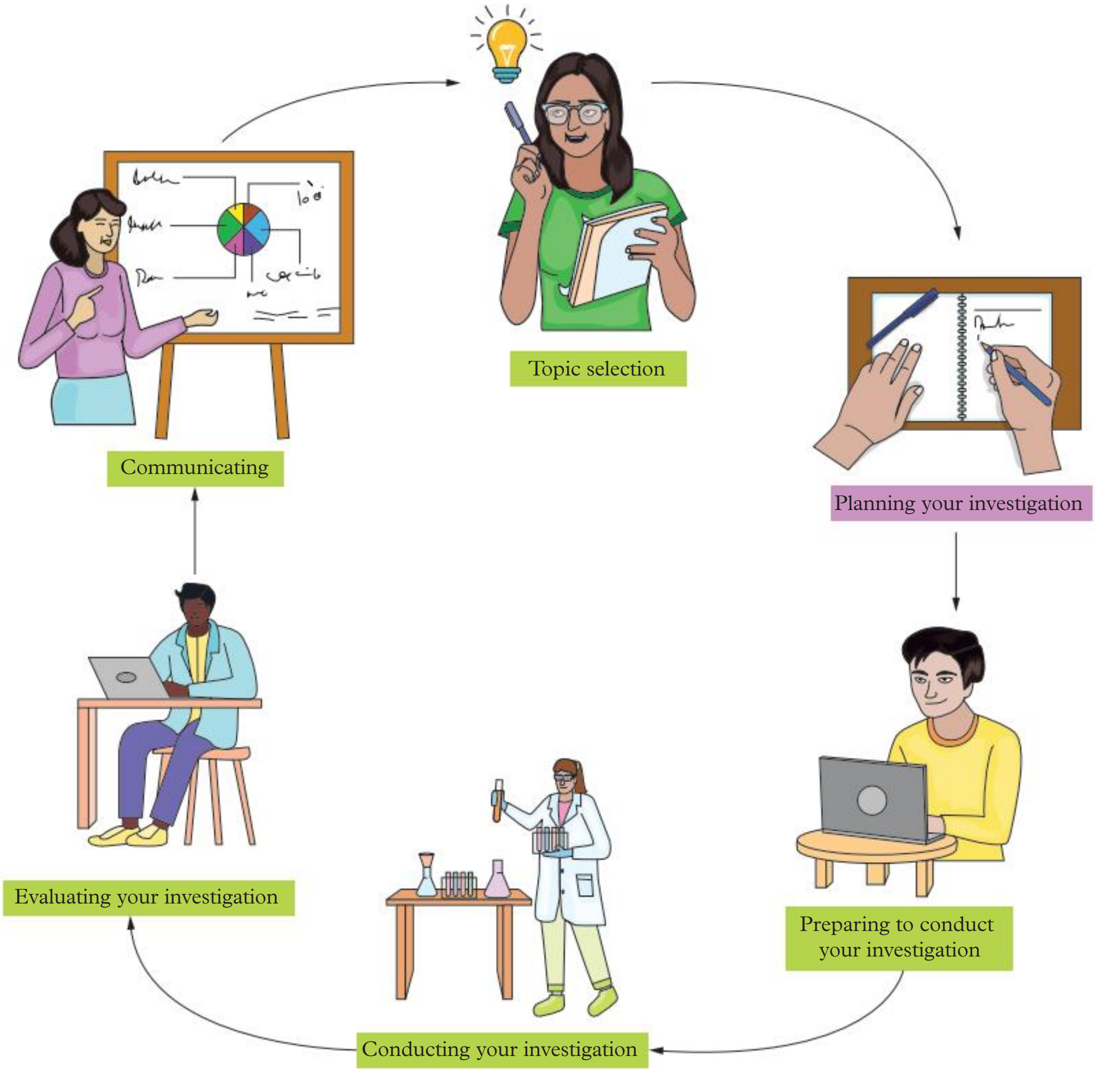


FIGURE 1 The scientific investigation process – this topic is about planning your investigation.

Defining the variables

Before you can write a hypothesis and method for your investigation, you must define your variables. The three types of variables to consider are the:

- independent variable (IV) – the thing you change
- dependent variable (DV) – the thing that you measure
- controlled variable(s) (CV) – the thing(s) you control (kept constant throughout the experiment).

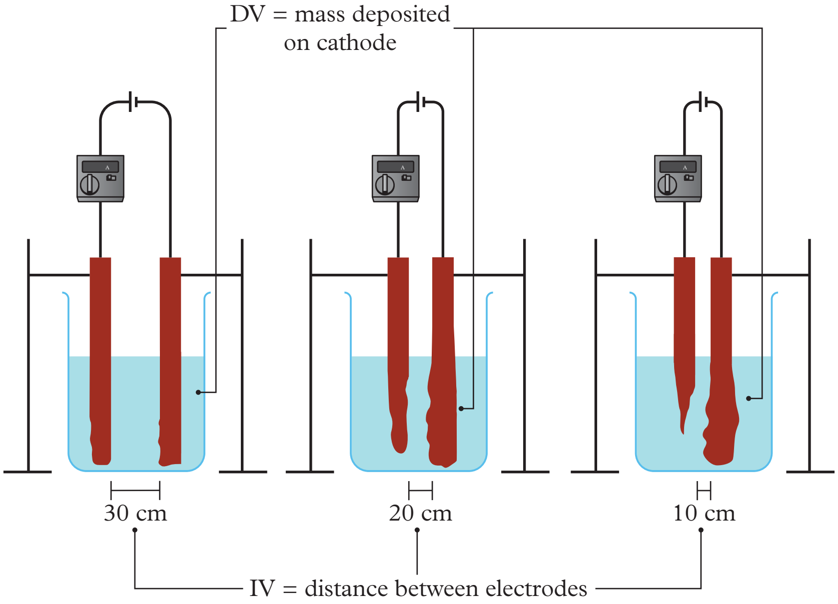


FIGURE 2 When testing the effect of distance between electrodes on the mass of metal deposited onto the cathode, IV is the distance between the electrodes, DV is the mass deposited at the cathode, and controlled variables could include material and size of the electrodes, volume and concentration of the electrolyte solution, and the duration that electricity is supplied to the cell.

Writing a testable hypothesis

Chapter 1 Chemistry toolkit showed you how to write a hypothesis. Now it's time to put this skill into practice. To write a hypothesis, you can use an 'If ..., then ... because ...' statement to make sure you cover all the elements you need.

If	Then	Because
If the **independent variable** is *[changed]*	then the **dependent variable** will *[change]*	because of **scientific reasoning**.
E.g. increased, decreased	*E.g. increase/decrease the amount/ rate/height/weight/number*	*A possible explanation for the relationship between the IV and DV*

FIGURE 3 A summary of how to write a testable hypothesis

Research tip

To remember the IV and DV roles and how you would graph them you can use DRY MIX:
DRY – **D**ependent (**R**esponding) **Y**-axis
MIX – **I**ndependent (**M**anipulated) **X**-axis

A useful hypothesis is a testable statement that often includes a prediction. In some instances, a research question may not lend itself to having an accompanying hypothesis. In such cases, students should work directly with their research questions. If this applies to your research question, check with your teacher before moving on, because some criteria might have marks allocated for a testable hypothesis and you may need to change your experiment.

Worked example 13.2 shows you how to write a hypothesis for the experiment in Figure 2.

13.2 WORKED EXAMPLE

WRITING A TESTABLE HYPOTHESIS FOR A RESEARCH QUESTION

Write a hypothesis for the research question below.

Research question: Does the distance between the electrodes in electroplating affect the amount of metal deposited onto the cathode?

Think	Do
Step 1: Define your variables.	IV: Distance between the electrodes DV: Mass of metal at the cathode Controlled variables: • Volume of electrolyte • Concentration of electrolyte • Size and material of electrodes • Time • Voltage • Size of electrodes (height/length) • Equipment – power pack, ammeter, glass container
Step 2: Write your hypothesis.	Hypothesis: *If* the distance between the electodes is decreased, *then* the mass of copper deposited onto the cathode will increase *because* as the distance between electrodes decreases, resistance between electrodes also decreases. This increases the current and subsequently the charge, which is directly proportional to mass. Therefore, the mass will increase.

Choosing a methodology

You can approach your scientific question in a number of ways. How you choose to approach your investigation is called your methodology. Different scientific methodologies include:

- controlled experiment
- fieldwork
- modelling
- simulation
- case study
- classification and identification
- product, process, or system development.

Information on each of these methodologies is provided in Chapter 1 Chemistry toolkit. The flow chart in Figure 4 can help you to choose the methodology best suited to your scientific question.

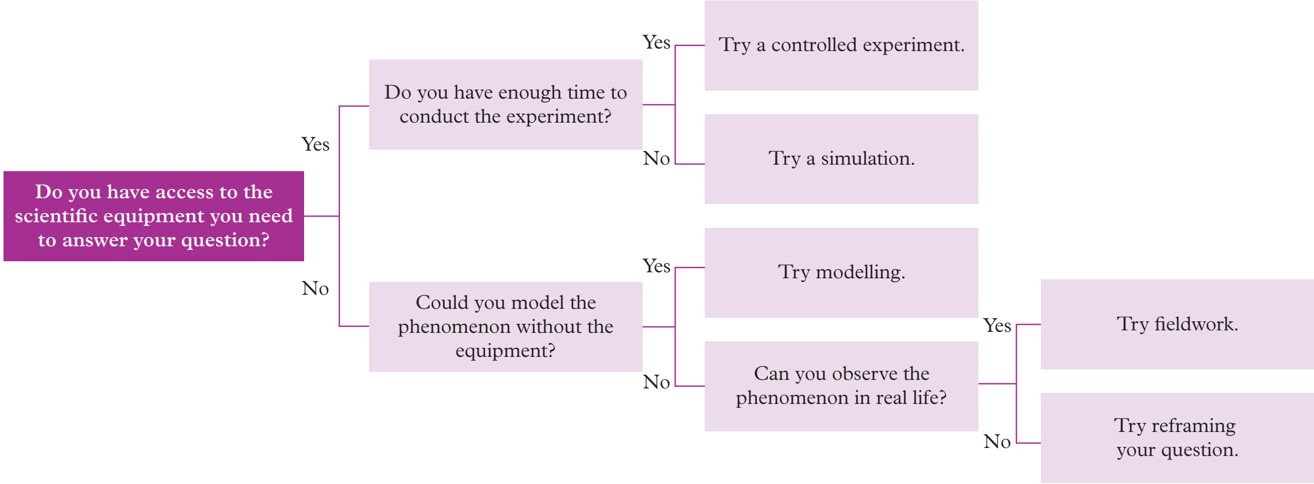


FIGURE 4 A flow chart to assist with selecting your scientific methodology

Generating primary data

For this investigation, you will need to generate your own primary data. There are two types of primary data that you can collect:

- **Quantitative data –** numerical data that is counted or measured
- **Qualitative data –** descriptions or categories based on observations

Tests you can run to generate primary data are summarised in Table 1.

FIGURE 5 Numeric values, such as time, are quantitative data.

TABLE 1 Approaches to collecting primary data

What kind of data do you collect?	Test	How do you do it?
Quantitative data	pH	pH test strips Universal indicator pH meter
	Mass	Gravimetric analysis Weighing mass change with an analytical balance
	Temperature	Using a thermometer
	Energy content	By determining specific heat capacity Calorimetry
	Rate	Measuring change in mass over time Measuring change in concentration over time
	Voltage and current	Using a multimeter, voltmeter or ammeter
Qualitative data	Reaction completion	Observing colour change in a reaction
	Properties of biodiesels	Observing complete or incomplete combustion Observing gelling with change in temperature
	Production of gases	Looking for bubbling Using the pop test for hydrogen gas
	Presence of CO_2	Adding limewater and observing results

Writing your method

For detailed information on how to write up your method, refer to Chapter 1 Chemistry toolkit. Key things to remember when writing a method are to:

- include detailed sequential steps:
  - ensure the steps are in correct order
  - number each of your steps
- ensure that your controlled and independent variables are appropriately manipulated:
  - the independent variable is the thing you change, so that change should be written in the method correctly
  - the controlled variables are the things you need to control during the experiment; providing detail on this is important for replication (see below)
- include appropriate measurements and concentrations with accurate use of units:
  - include the quantity of each item
  - include correct units
  - include the correct equipment
- write the method in past tense.

TO-DO LIST

- Choose a methodology for your investigation and justify why you have chosen this as the best way in which to answer your scientific question.
- Identify the independent and dependent variables that you will be investigating.
- Identify which variables you will control for your investigation.
- Write a testable hypothesis for your investigation, using the 'If …, then … because …' style.
- Write a succinct and detailed method for your investigation. (Remember to include enough detail so someone else could conduct your investigation for you – aim for repeatability and reproducibility.)

13.3

Preparing to conduct your investigation

KEY GOALS

In this topic, you will:

- complete a risk assessment
- consider the ethics around your experiment.

In this topic, we will look at how to prepare to conduct your scientific investigation. Preparing to conduct your investigation includes:

- conducting a risk assessment
- considering the ethics of your investigation.

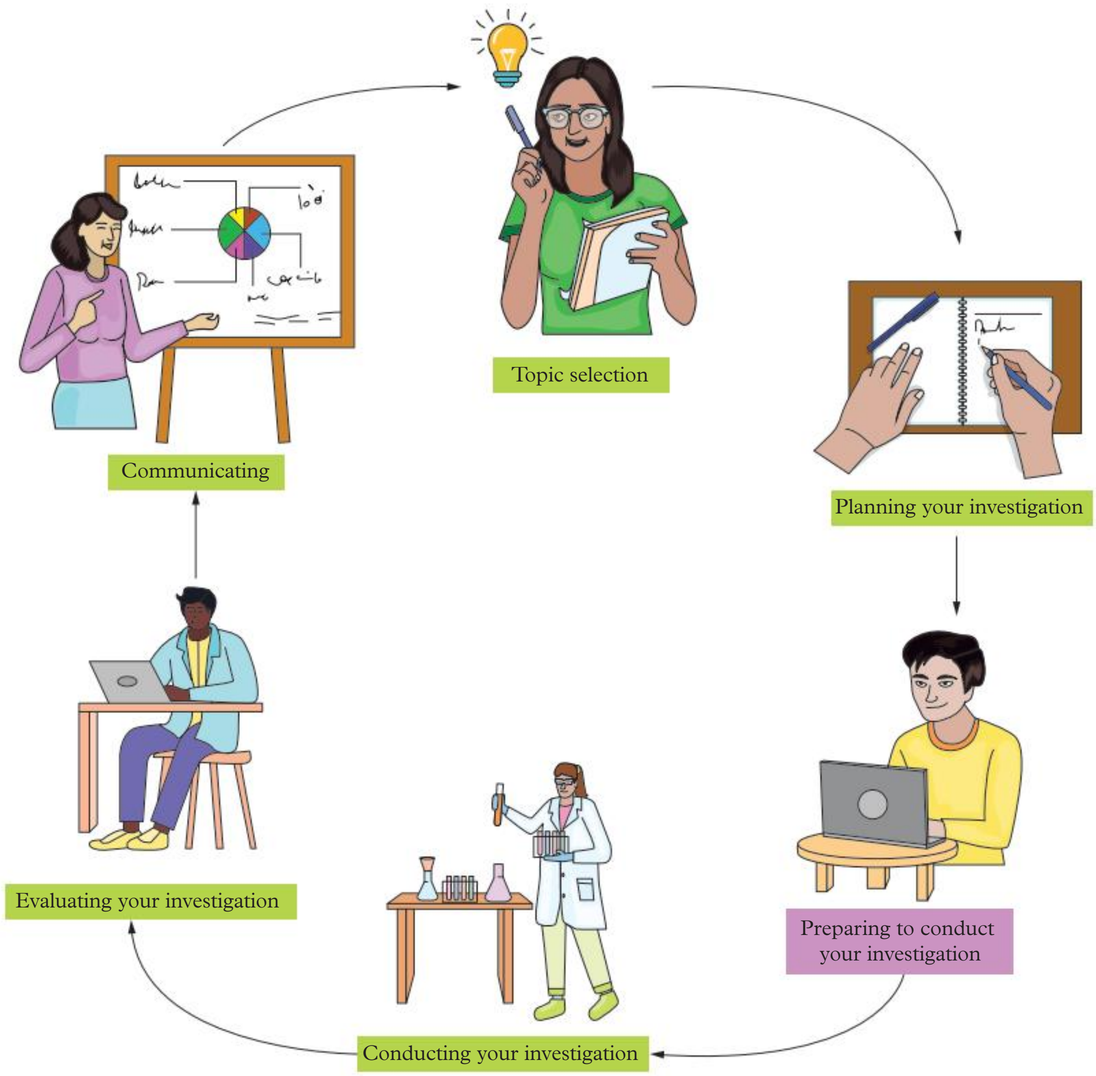


FIGURE 1 The scientific investigation process – this topic is about preparing to conduct your investigation.

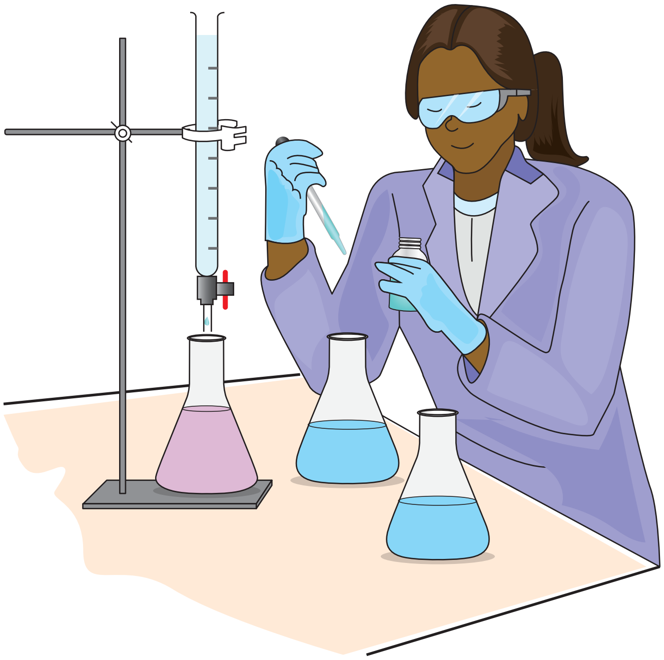


FIGURE 2 Personal protective equipment (PPE), including safety glasses and lab coats, should be worn to protect you during an experiment.

Conducting a risk assessment

When conducting an experiment, there are many health and safety considerations you need to take to protect yourself and those around you in the laboratory. Before you conduct an experiment, you must conduct a risk assessment. Risk assessments help you identify all the potential risks associated with your materials, chemicals and health and safety.

The instructions for conducting a risk assessment for an experiment are outlined in Chapter 1 Chemistry toolkit (page 17). A risk assessment template is also provided in your Student obook pro for you to download and print.

Ethical considerations

For your scientific investigation, you will need to generate primary data. This means when writing and planning your investigation, you will need to consider any potential ethical considerations. This includes considering the impact of your research on:

- the greater community
- the environment
- living organisms and non-living things.

Other ethical factors to be aware of when preparing for your investigation include:

- the sourcing of materials
- disposal of chemicals and equipment
- the green chemistry principles
- the United Nations Sustainable Development Goals.

You are expected to apply integrity when recording and reporting results and data. This means you must not falsify results and you must acknowledge prior research. It is also important to consider the impact of your results and how they may be used in the future by others.

TO-DO LIST

- Conduct a risk assessment for your investigation.
- Evaluate the ethics of your investigation.

13.4

Conducting your investigation

KEY GOALS

In this topic, you will:

- create a data table in which you will collect and record your data
- prepare to conduct your experiment.

In this topic, you we will look at conducting your scientific investigation. This will include:

- generating and recording data
- conducting the experiment.

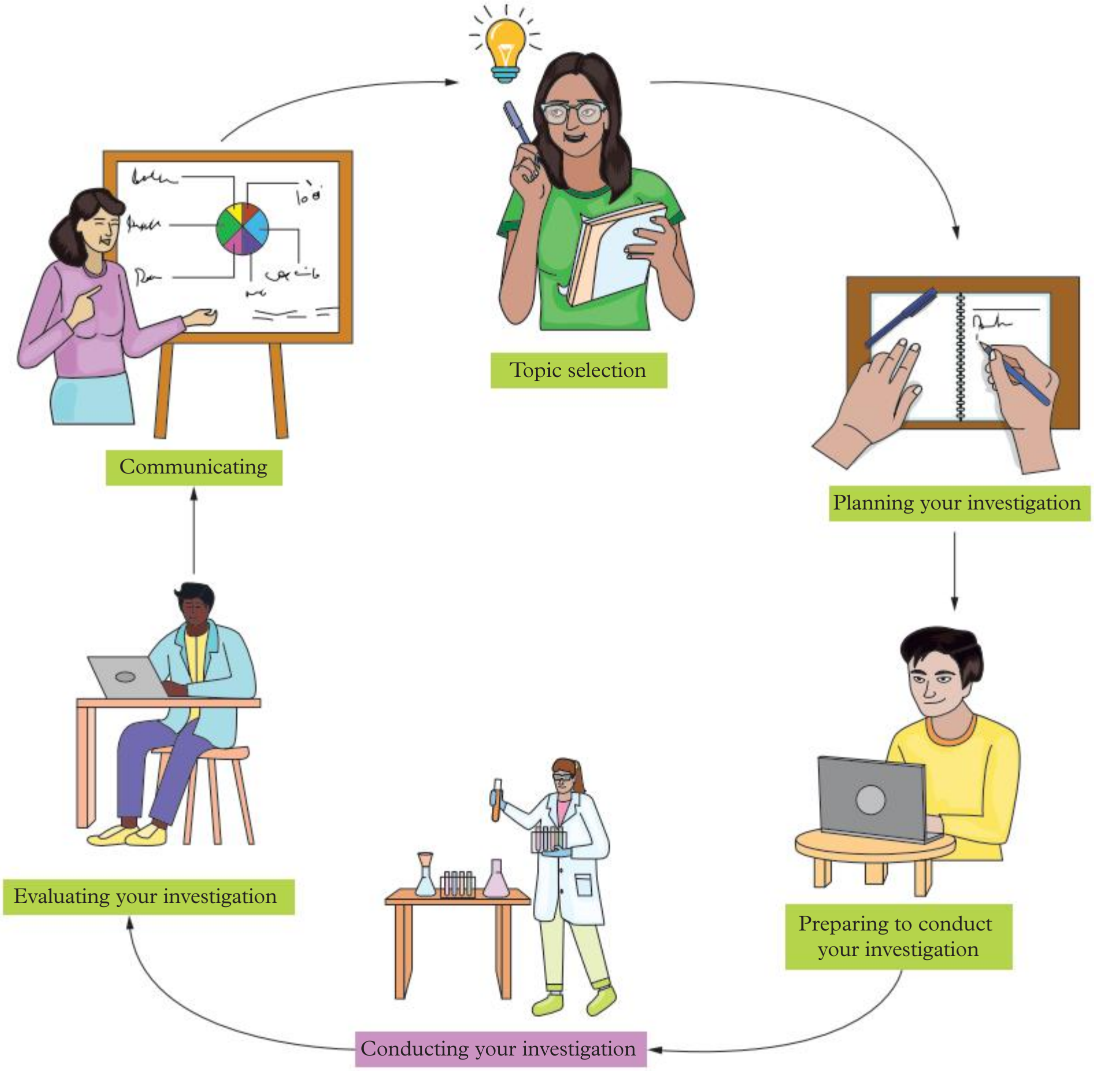


FIGURE 1 The scientific investigation process – this topic is about conducting your scientific investigation.

Recording results, data and observations

Establishing good data-recording habits will help you greatly when it comes to writing your discussion, and evaluating your errors. Data collection does not just include recording numbers and data, but also recording your observations.

Figure 2 shows an example of an annotated and complete table for recording results, data and observations for the practical investigation 'Does the distance between the electrodes in electroplating affect the amount of metal deposited onto the cathode?'

Trial number	**IV: Distance between electrodes (cm)**	**DV: Mass of the metal at cathode (g)**				**Current (amps)**	**Observations**
		Initial	**Final**	**Δ mass**	**Average mass**		
1	30	0.457	0.466	0.009	0.0085	0.26	*t* = 120 s; cathode removed first
2		0.491	0.499	0.008		0.24	*t* = 120 s; cathode removed first
3		0.466	0.480	0.014		0.25	Cell was left on while electrodes were being removed – extra time received Personal error result not included in average.
1	20	0.448	0.463	0.015	0.0153	0.28	*t* = 120 s; cathode removed first
2		0.483	0.498	0.015		0.30	*t* = 120 s; cathode removed first
3		0.503	0.519	0.016		0.30	*t* = 120 s; anode accidentally removed first
1	10	0.499	0.522	0.023	0.0225	0.38	*t* = 120 s; cathode removed first
2		0.516	0.534	0.018		0.36	*t* = 120 s; cathode removed first Some mass fell off electrode when removing from cell. Personal error not included in average.
3		0.483	0.505	0.022		0.38	*t* = 120 s; cathode removed first

Headings were well labelled, including units.

All masses were recorded (initial and final). Change in mass and average mass were calculated.

Current was measured so theoretical results could be checked and the hypothesis could be confirmed/rejected.

Time was controlled and accounted for.

Errors made were noted and not included in results.

The student noted the order of cathode/anode being removed.

Three trials were done so that averages could be calculated and any errors removed.

FIGURE 2 An annotated and completed table for recording results, data and observations when testing the effect of soil pH on soybean growth

There may be a considerable amount of time between when you first conduct the experiment and when you write it up. Good records of data, results and observation will be extremely helpful when you need to write because they will refresh the details of the experiment. Record your results in your logbook because these will need to be verified for authenticity by your teacher.

Research tip

When conducting an experiment, scientists constantly take notes and annotate changes to the method. This ensures the method can be reproduced by others in the future.

Conducting your experiment

When conducting your experiment, you should remember the following:

- Be safe at all times.
- Consider the safety of those around you.
- Follow your method or make annotations in your logbook when you change anything.
- Record all results at the time of taking them.
- Record any errors or things that seem out of place.
- Record observations and comparisons between tests.
- Check before disposing of experiment equipment, chemicals and materials.
- If you have any questions, always ask your teacher.

FIGURE 3 When setting up and conducting your research make sure you are following the correct safety procedures.

TO-DO LIST

- Set up a logbook so that you can record all results and observations as you conduct the investigation.
- Check with your teacher to make sure you are ready to begin generating data.

13.5

Evaluating your investigation

KEY GOALS

In this topic, you will:

- understand different ways to display your results
- evaluate your results
- identify and discuss any errors
- write a discussion about your investigation.

In this topic, you will look at evaluating your scientific investigation. This includes:

- displaying your data
- writing a discussion
  - analysing your results
  - evaluating your errors
  - modifications to the method
- writing a conclusion.

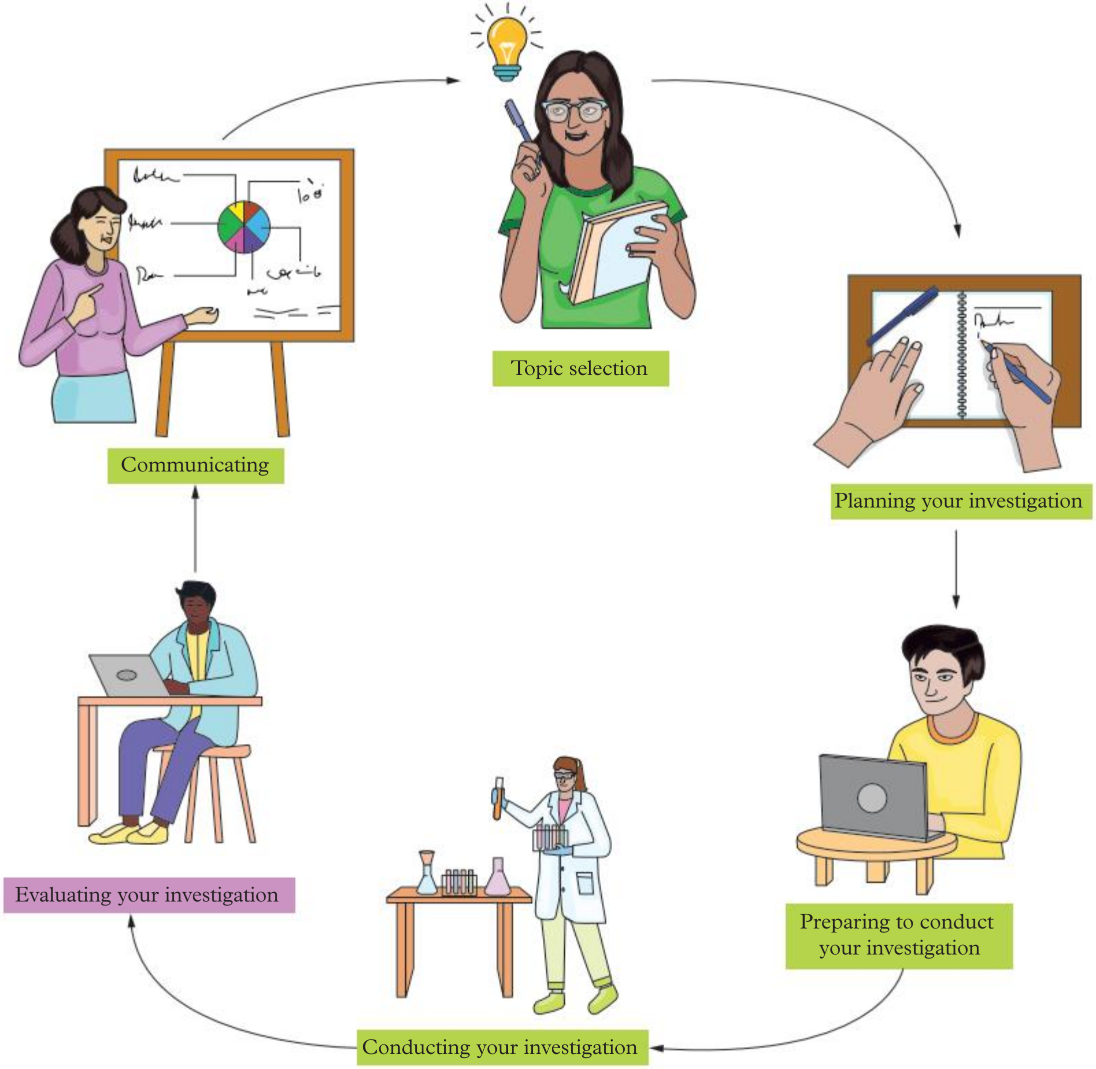


FIGURE 1 The scientific investigation process – this topic is about evaluating your investigation.

Displaying your data

After conducting your experiment and recording your data and observations, you will need to select the right way to display your data so that others can get a quick snapshot of your results. Your data presentation should also outline any trends that may exist between your variables.

The way in which you display your data is important; some graphs cannot be used for a particular purpose. For example, you cannot use pie charts to compare data. The flow chart in Figure 2 summarises different methods of displaying data depending on how you wish to use the information.

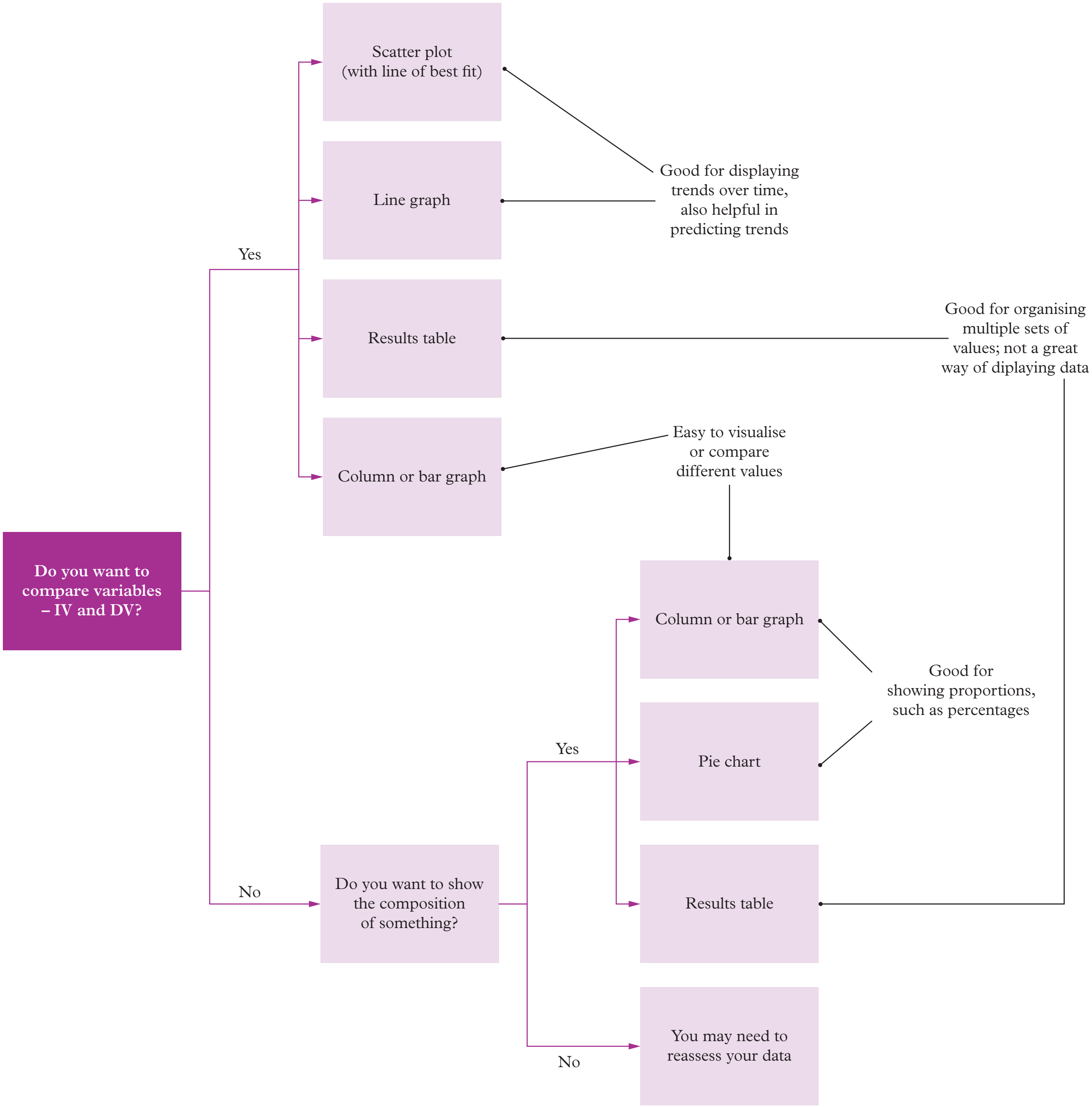


FIGURE 2 A flow chart to assist your decision on how you will display your data

Remember that what you are trying to show the people reading or assessing your investigation is the relationship between variables, or the trend or pattern from your results.

13.5A WORKED EXAMPLE

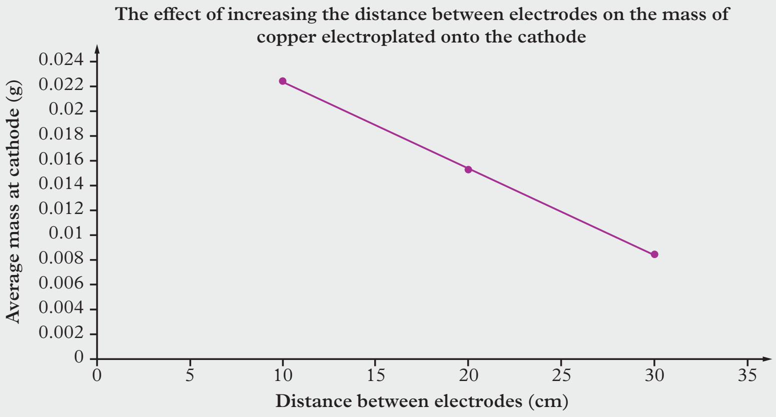
DISPLAYING DATA

The data in Table 1 was collected in an investigation exploring the research question: 'Does the distance between the electrodes in electroplating affect the amount of metal deposited onto the cathode?'

TABLE 1 A results table of the distance between electrodes and the change in mass at the cathode

Distance between the electrodes (cm)	Increase in mass at cathode (average) (g)
30	0.0085
20	0.0153
10	0.0225

Select and display the results from the table on the most appropriate graph.

Think	Do
Step 1: Use the flow chart in Figure 2 to determine which graph type best suits the data.	The data has variables to compare, so we could use a line graph or a bar chart. However, there is also a distinct relationship between the mass at the cathode and the distance between the electrodes. This relationship is more evident when displayed through a line graph instead of a bar chart.
Step 2: Graph the data in the table.	Line graph of the distance between the electrodes and the mass electroplated at the cathode: 

Research tip

Play around with different graph styles if you are having trouble deciding. There are graphing programs that will give you a lot of different types of graphs; if in doubt, ask someone to tell you what your graph says to them.

Writing a discussion

The information you need to provide in your discussion depends on what the assessment criteria for your investigation says to include. For the most part, it will include:

- analysis of your results
- evaluation of your errors
- modifications of the method.

Research tip

If you are drawing graphs that have an x- and y-axis, then the IV is on the x-axis and the DV is on the y-axis.

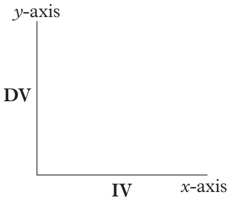


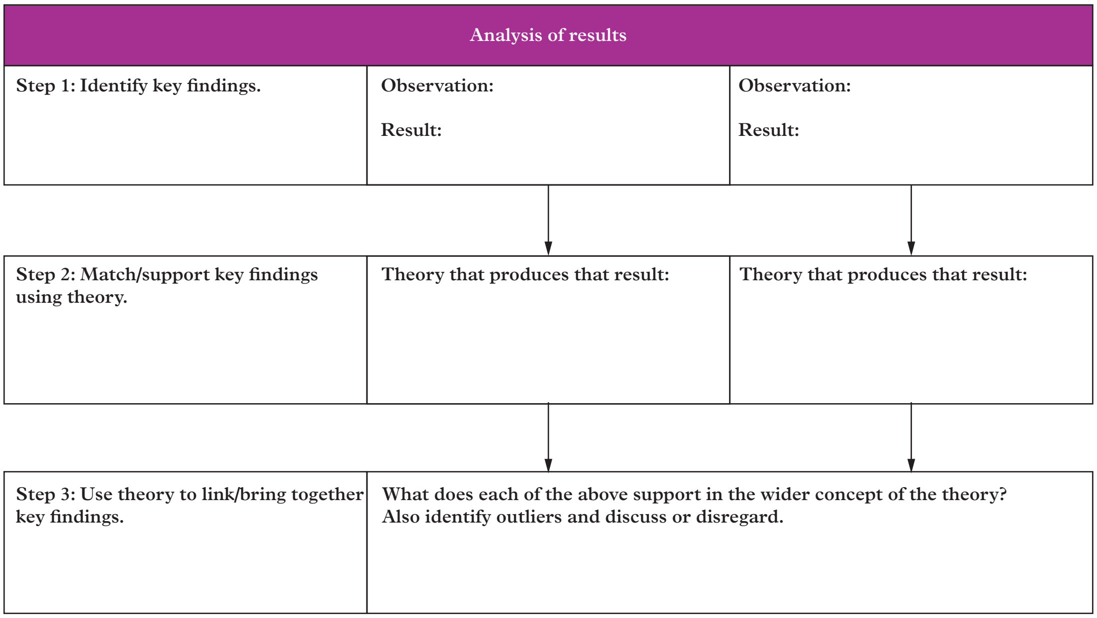
FIGURE 3 How to graph your IV and DV

Analysis of your results

After displaying your data to show your results and trends, it is time to analyse and evaluate your results. To do this, there are a few questions you need to ask and answer:

- What trend can you see from the data?
- How does the trend relate to the theory?

One approach to addressing these questions is to simply go through each of the questions and answer them. If you need more guidance, another approach is to use the discussion flow chart in Figure 4. Figure 4 can help you understand how to answer each question by unpacking it, and format your responses so they can be included in your discussion.



Analysis of results		
Step 1: Identify key findings.	Observation: Result:	Observation: Result:
Step 2: Match/support key findings using theory.	Theory that produces that result:	Theory that produces that result:
Step 3: Use theory to link/bring together key findings.	What does each of the above support in the wider concept of the theory? Also identify outliers and discuss or disregard.	

FIGURE 4 A flow chart for analysing results for discussion

For example, Figure 5 on the next page shows annotated results from an investigation on the effect of the distance between electrodes on the change in mass at the cathode. You will need to use research and your understanding to link and support the findings from your experiment. You may use information you have already gathered from the topic selection phase. or conduct more research to inform and support your data and trends.

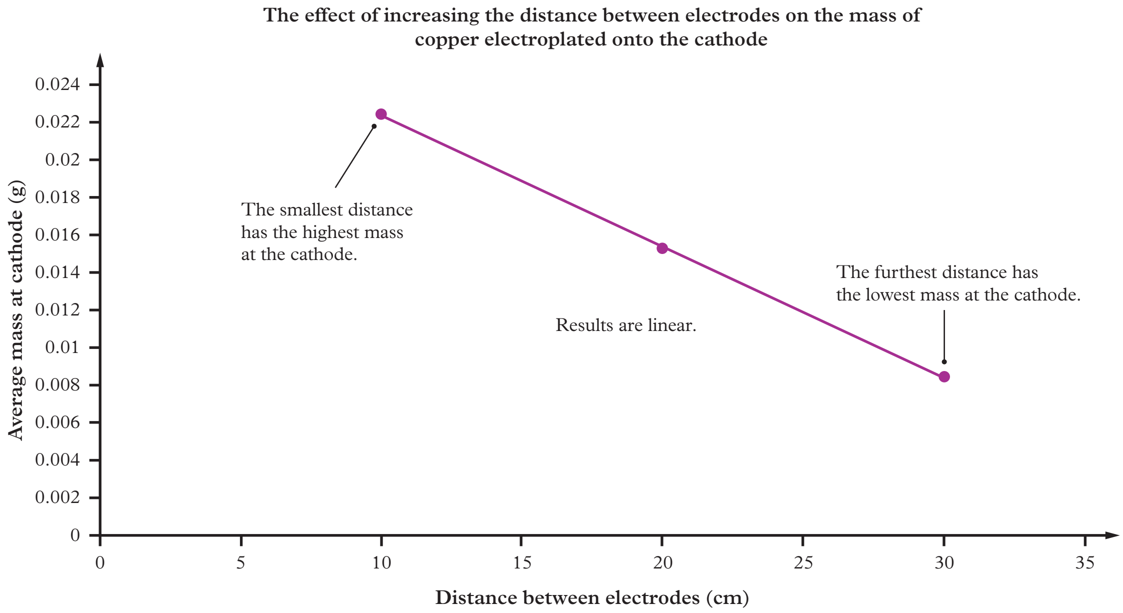


FIGURE 5 Annotation of the displayed results from the investigation of the distance between electrodes and the mass of copper electroplated

You may also need to gather more information if you find that your data or trend is not supported by the theory, or that your hypothesis is not supported. Figure 6 shows a worked example of how you can use the flow chart for analysis of data to inform your discussion.

Analysis of results		
Step 1: Identify key findings.	**Observation:** The highest mass was produced at 10 cm, and the lowest mass was produced at 30 cm. **Result:** The trend was that as the distance between the electrodes decreased, the mass at the cathode increased.	**Observation:** The current was lowest at 30 cm, then at 20 cm and 10 cm. **Result:** The current increased as the distance between the electrodes decreased.
Step 2: Match/support key findings using theory.	**Theory that produces that result:** As the distance between the electrodes decreased, the space between the electrodes had less resistance, so more metal was able to transfer onto the cathode. The decrease in resistance means the current was increased in the cell. Therefore, Q increased and the mass at the cathode increased.	**Theory that produces that result:** As distance between the electrodes decreased, and the current increased, the resistance decreased: $R = V/I$ Resistance 30 cm = 6 V/0.25 A = 24 Ω 20 cm = 6 V/0.29 A = 21 Ω 10 cm = 6 V/0.38 A = 15 Ω There is less space and fewer ions between the electrodes. This caused the current in the cell to increase. Therefore, Q increased and the mass at the cathode increased.
Step 3: Use theory to link/bring together key findings.	**What do steps 1 and 2 support in the wider concept of the theory? Also identify outliers and discuss or disregard.** No outliers – but only one set of results for each: • There are no outliers, but personal errors were not included in the averages. • As the distance between the electrodes decreased, the resistance between the electrodes decreased. • The resistance decreasing caused the current required by the cell to increase, as $R = V/I$ and voltage was kept constant. • An increase in current means that more electricity (Q) will be required, as $Q = I \times t$. • As Q is increased, more mass will be plated onto the cathode, as mass is directly proportional to charge ($m \propto M$).	

FIGURE 6 An example of how the analysis of data flow chart can be used

Evaluating errors

Chapter 1 Chemistry toolkit explains types of errors, accuracy, precision and reproducibility, repeatability and reliability. You can refresh your memory of each of these by reading pages 24–25 in Chapter 1. Figure 7 shows how errors impact each of these factors.

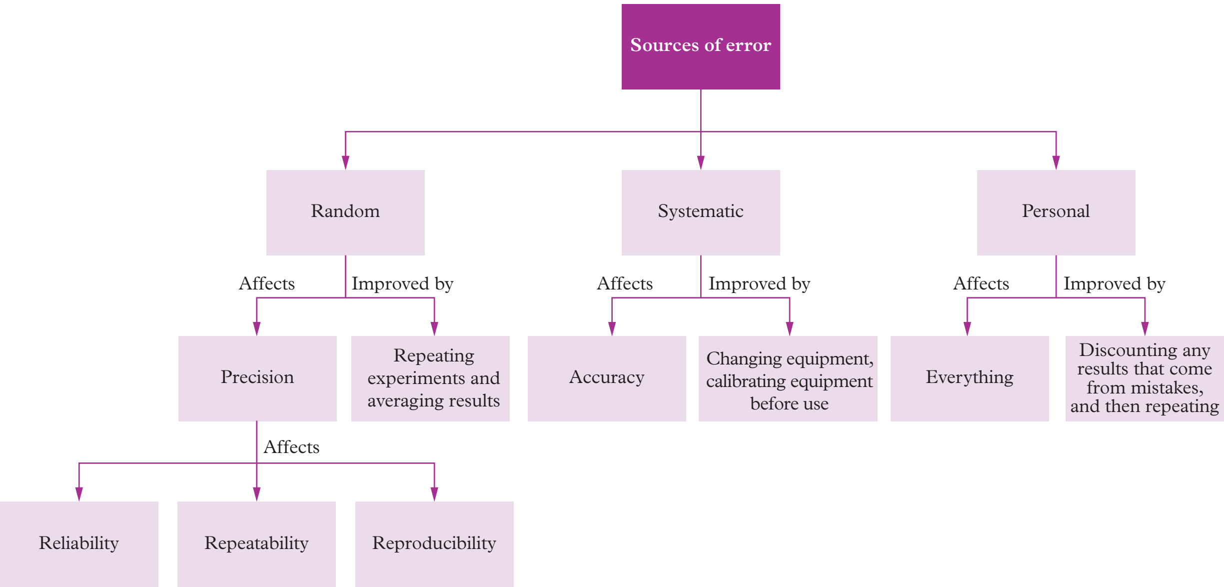
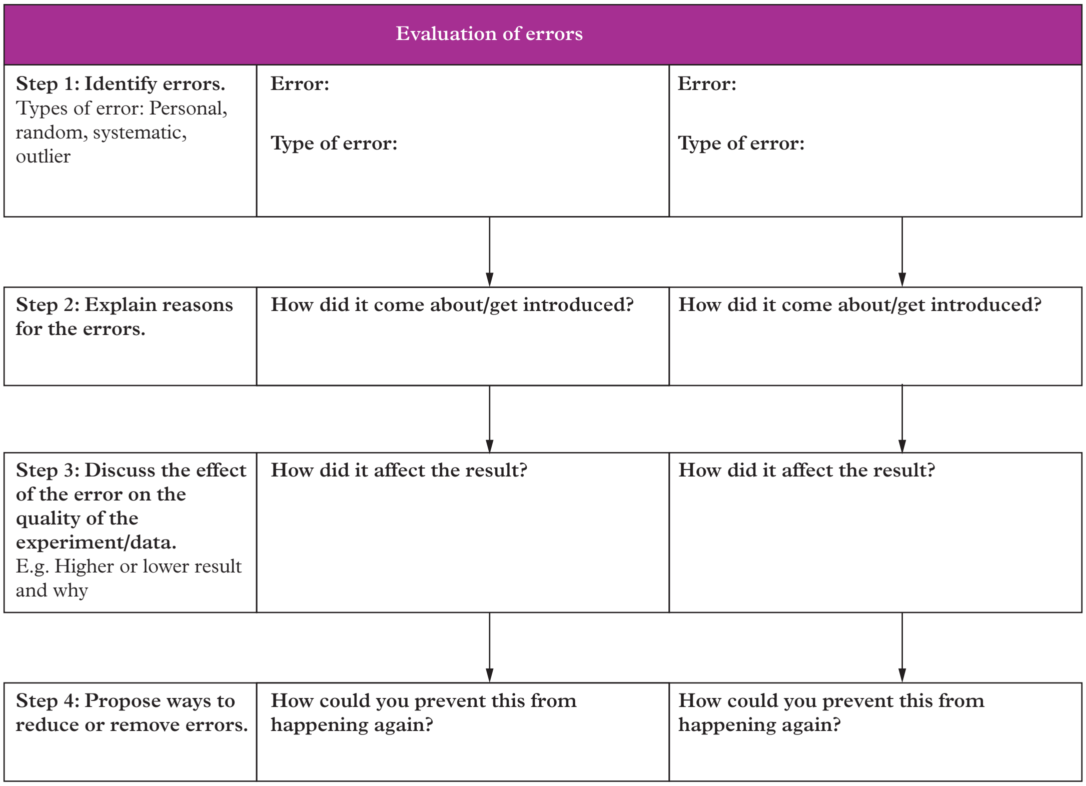


FIGURE 7 Errors and their effects and how they can be improved during an experiment

It is important that you evaluate your errors in your discussion. Figure 8 is a flow chart that shows how you can evaluate your errors and format your responses to be included in your discussion.



Evaluation of errors		
Step 1: Identify errors. Types of error: Personal, random, systematic, outlier	**Error:** **Type of error:**	**Error:** **Type of error:**
Step 2: Explain reasons for the errors.	**How did it come about/get introduced?**	**How did it come about/get introduced?**
Step 3: Discuss the effect of the error on the quality of the experiment/data. E.g. Higher or lower result and why	**How did it affect the result?**	**How did it affect the result?**
Step 4: Propose ways to reduce or remove errors.	**How could you prevent this from happening again?**	**How could you prevent this from happening again?**

FIGURE 8 A flow chart for evaluating errors for discussion

An example how to use the error evaluation flow chart is shown in Figure 9.

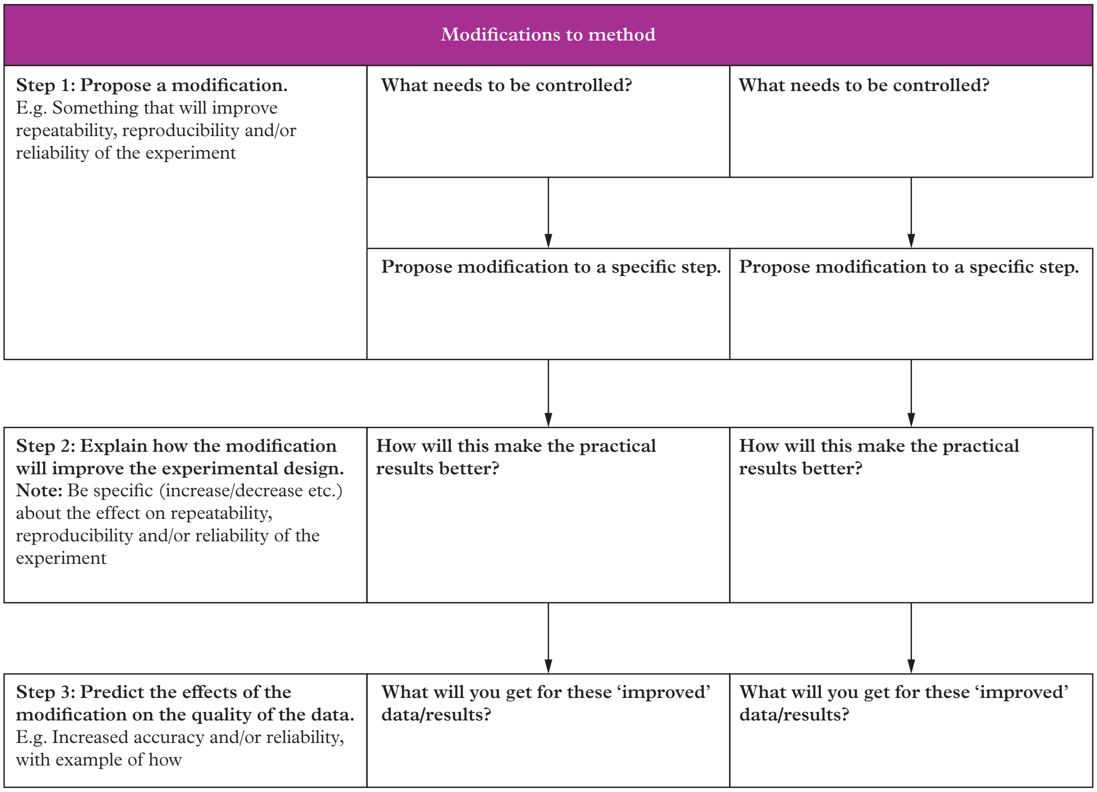
Evaluation of errors		
Step 1: Identify errors. Types of error: Personal, random, systematic, outlier	**Error:** There was a higher change in mass in trial 3 for 30 cm. **Type of error:** Personal	**Error:** There was a lower change in mass in trial 2 for 10 cm. **Type of error:** Personal
Step 2: Explain reasons for the errors.	**How did it come about/get introduced?** The cell was left on while electrodes were being removed, leading to more time in the cell. Therefore, a higher mass was plated.	**How did it come about/get introduced?** Some mass fell off as the electrode was being removed from the electroplating cell. The recorded mass was lower.
Step 3: Discuss the effect of the error on the quality of the experiment/data. E.g. Higher or lower result and why	**How did it affect the result?** The result was excluded from the average calculation as it was a personal error. If it were included, it would have increased the average mass and the result would not be linear.	**How did it affect the result?** The result was excluded from the average calculation as it was a personal error. If it were included, it would have decreased the average mass and the result would not be linear.
Step 4: Propose ways to reduce or remove errors.	**How could you prevent this from happening again?** Make sure the cell power is turned off as soon as the stopwatch hits 120 s. Repeat the trial so that there are three trials to average.	**How could you prevent this from happening again?** Take care when removing the electrodes from the cell. Repeat the trial so that three cells can be averaged.

FIGURE 9 Evaluating errors using the discussion flow chart for the investigation of soil and plant growth

Modifying the method

Any errors that occurred during your investigation will affect the repeatability, reproducibility and reliability of your experiment. You need to consider your controlled variables during your experiment and when you are evaluating your errors. Figure 11 shows a flow chart that can help you determine modifications you can make to your method.

FIGURE 10 Be sure to keep track of any errors that occur during your experiment.



Modifications to method

Step 1: Propose a modification. E.g. Something that will improve repeatability, reproducibility and/or reliability of the experiment

What needs to be controlled?

What needs to be controlled?

Propose modification to a specific step.

Propose modification to a specific step.

Step 2: Explain how the modification will improve the experimental design. **Note:** Be specific (increase/decrease etc.) about the effect on repeatability, reproducibility and/or reliability of the experiment

How will this make the practical results better?

How will this make the practical results better?

Step 3: Predict the effects of the modification on the quality of the data. E.g. Increased accuracy and/or reliability, with example of how

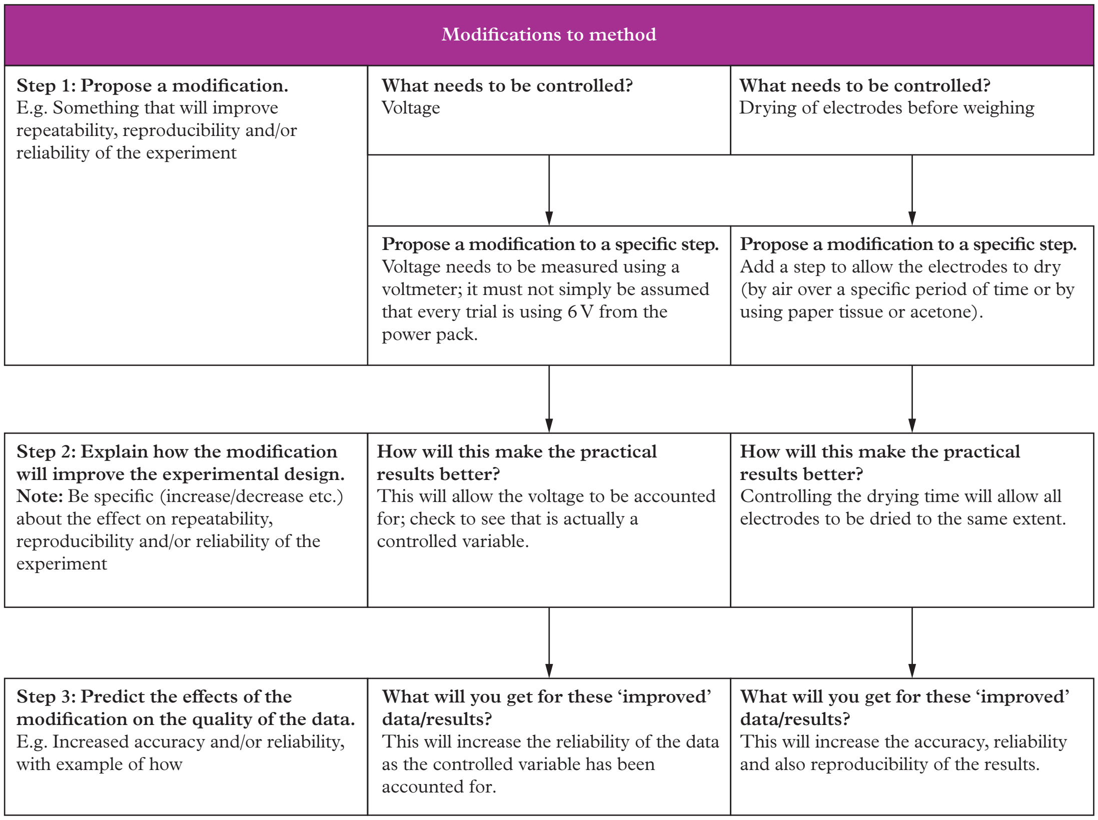
What will you get for these 'improved' data/results?

What will you get for these 'improved' data/results?

FIGURE 11 Modifications to method flow chart for discussion

FIGURE 12 Think about how you could modify your method to reduce potential errors. For example, repeating experiments and averaging results reduces random error.

An example of how to use the modifications to method flow chart to add to your discussion is given in Figure 13.



Modifications to method		
Step 1: Propose a modification. E.g. Something that will improve repeatability, reproducibility and/or reliability of the experiment	**What needs to be controlled?** Voltage	**What needs to be controlled?** Drying of electrodes before weighing
	Propose a modification to a specific step. Voltage needs to be measured using a voltmeter; it must not simply be assumed that every trial is using 6 V from the power pack.	**Propose a modification to a specific step.** Add a step to allow the electrodes to dry (by air over a specific period of time or by using paper tissue or acetone).
Step 2: Explain how the modification will improve the experimental design. **Note:** Be specific (increase/decrease etc.) about the effect on repeatability, reproducibility and/or reliability of the experiment	**How will this make the practical results better?** This will allow the voltage to be accounted for; check to see that is actually a controlled variable.	**How will this make the practical results better?** Controlling the drying time will allow all electrodes to be dried to the same extent.
Step 3: Predict the effects of the modification on the quality of the data. E.g. Increased accuracy and/or reliability, with example of how	**What will you get for these 'improved' data/results?** This will increase the reliability of the data as the controlled variable has been accounted for.	**What will you get for these 'improved' data/results?** This will increase the accuracy, reliability and also reproducibility of the results.

FIGURE 13 An example of the modifications to method flow chart for the investigation regarding soil and plant growth

Conclusion

The conclusion is the end to your poster, report or presentation, a wrap-up of everything you did. You should check the assessment criteria for your investigation to write your conclusion, but conclusions often involve answering the following questions.

- What were your key findings?
- Was your hypothesis supported or disproved?
- Were there any limitations in your experiment?
- How can you overcome those limitations?

Was your hypothesis supported or disproved?

It is important to link all aspects of your poster to the key idea – the question, or hypothesis. In your conclusion, you should first summarise the key findings, then link them straight to the hypothesis and explain if they supported your hypothesis or disproved it.

Limitations

In your conclusion, you need to consider the limitations in your experiment. Check if these are assessed in your conclusion or in your discussion. Limitations are things that affect the overall research and conclusion that you are making. For example, if you were testing the corrosion of metals, and only used zinc to test it, then you cannot conclude that all metals will act this way. The final step in a conclusion is to propose a way that you could overcome these limitations.

13.5B WORKED EXAMPLE

WRITING A CONCLUSION

Write a conclusion for the investigation on the effect of the distance between electrodes and the change in mass at the cathode.

Think	Do
Step 1: Summarise your key findings.	As the distance between the electrodes decreased, the mass of metal electroplated on the cathode increased.
Step 2: Link your key findings to your hypothesis (has it been supported or disproved).	The initial hypothesis was 'If the distance between the electrodes is decreased, then the mass of copper deposited onto the cathode will increase.' The hypothesis was supported because findings showed that the 30 cm distance produced the lowest mass, while the 10 cm distance produced the highest mass.
Step 3: Identify any limitations.	One limitation was that only three distances were tested. While a linear trend was produced, that may change with smaller or larger distances. Another limitation was that the investigation only tested copper and not other metals.
Step 4: Provide recommendations to overcome limitations.	Limitations can be overcome by testing a larger range of distances between electrodes. The investigation can also be expanded to include other electroplating cells like zinc and iron to see if the trend is always linear.

TO-DO LIST

- Find the best way to display your data to best represent your results.
- Analyse your results:
  - identify key findings
  - match/support key findings using theory
  - link key findings and theory.
- Evaluate your errors:
  - identify errors
  - explain reasons for errors
  - discuss effects of errors on quality of data
  - propose way to remove or reduce errors.
- Modify the method:
  - propose modifications to the method
  - explain how modifications will improve experiment
  - predict effect of modifications on quality of data.
- Identify whether your hypothesis was supported.
- Identify the limitations of your investigation design.
- Propose recommendations to overcome limitations.

13.6

Communicating your findings

KEY GOALS

In this topic, you will:

- select a format for your investigation
- understand where everything needs to go on your poster
- create a reference list and acknowledgements.

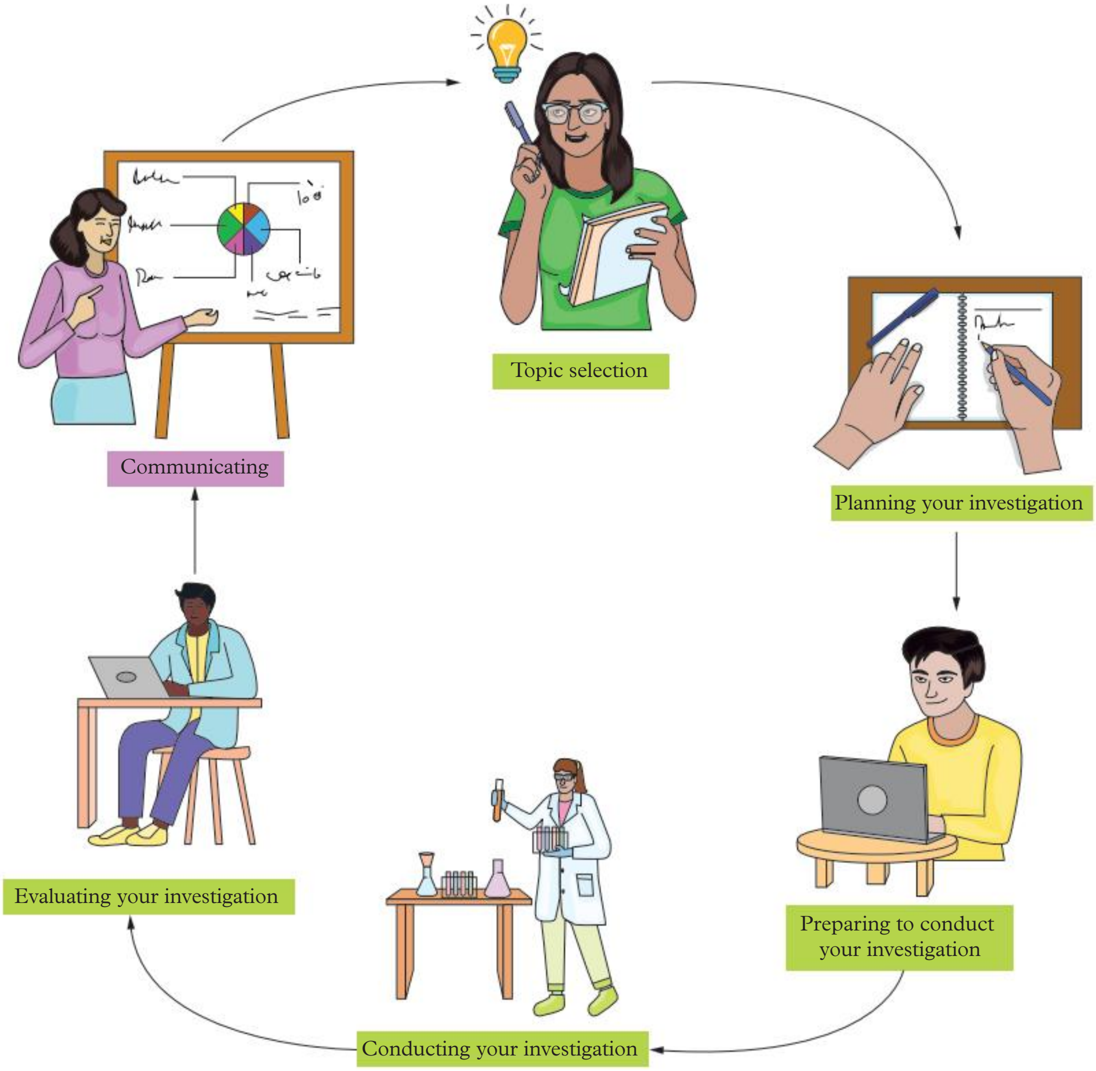


FIGURE 1 The scientific investigation process – this topic is about communicating your findings.

In this topic, we will look at the reporting phase of your scientific investigation. This will include:

- communicating your results as a poster
- using scientific terminology
- referencing
- acknowledgements.

Communicating your results as a poster

FIGURE 2 Scientific posters on display at a science conference

Unit 4 Area of Study 3 requires you to report your findings as a scientific poster. The poster format outlined in the VCAA study design is shown in Figure 3, which includes VCAA's suggestions for what should be under each heading. The poster may be produced electronically or in hard-copy format and must not exceed 600 words.

The 600-word limit does not include supporting text, such as:

- tables
- graphs
- image captions
- references
- acknowledgements.

<table>
<tr><th colspan="3">Title
Student name</th></tr>
<tr>
<td>Introduction
• Brief explanation or reason for undertaking the investigation, including a clear aim, a hypothesis and/or a prediction and relevant background chemical concepts
Methodology and methods
• Brief outline of the selected methodology used to address the investigation question
• Summary of data generation method(s) and data analysis method(s)
Results
• Presentation of generated data/evidence in appropriate format to illustrate trends, patterns and/or relationships</td>
<td>Communication statement reporting the key finding of the investigation as a one-sentence summary</td>
<td>Discussion
• Interpretation and evaluation of analysed primary data
• Identification of limitations in data and methods, and suggested improvements
• Cross-referencing of results to relevant chemical concepts
• Linking of results to investigation question and to the aim to explain whether the investigation data and findings support the hypothesis
Conclusion
• Conclusion that provides a response to the investigation question
• Identification of the extent to which the analysis has answered the investigation question, with no new information being introduced</td>
</tr>
<tr><td colspan="3">References and acknowledgements
• Referencing and acknowledgement of all quotations and sourced content relevant to the investigation</td></tr>
</table>

FIGURE 3 The VCAA scientific poster format

Source: *VCE Chemistry Study Design (2023–2027)* reproduced by permission © VCAA

Within the 600 words, you need to summarise everything that you have done throughout your investigation. This means it is important be careful with how many words you designate to each section.

The following tips can help you generate your scientific poster or report:

- **Choose your content wisely.** Don’t include fluff and things that don’t relate to your question; find one or two *key* points and stick with those throughout the poster.
- **Don’t waste words when there are only a small number of marks allocated.** Check the assessment criteria – if there is only one mark for something, then don’t waste 200 words trying to explain it.
- **Proofread your work.** Edit your poster to make sure there is no irrelevant information. After proofreading once, go back and proofread/edit again. Having a friend proofread it for you is also valuable.
- **Make your poster visually appealing.** Figure 4 shows the dos and don’ts of formatting your poster.

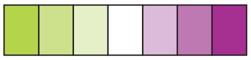
DO	DON’T
• Do use one colour scheme throughout your poster; e.g. 	• Don’t use more than two font types, because “MULTIPLE FONTS can **make your** poster *look messy*.”
• Do replace or break up large sections of text with figures; e.g. 	• Don’t use clashing or garish colours; e.g. 
• Do leave some blank space so that your poster is not too text heavy.	• Don’t use all capital letters; e.g. ‘THE HYPOTHESIS WAS REJECTED’

FIGURE 4 A summary of what to do and what not to do to make your poster visually appealing

Your other assessment tasks in Units 3 and 4 may require you to communicate scientific findings in another format, such as a report or a video. The same tips apply.

The communication statement

The centre of the poster will occupy 20–25% of the poster space and will be a one-sentence summary of the major finding of the investigation that answers the investigation question.

Because the communication statement is such a large area of your poster, when writing your statement, you should:

- keep it succinct and use just one sentence
- avoid using long, confusing words
- avoid phrasing it as a question
- make sure it relates to the question.

For example, a communication statement based on the sample investigation covered in this chapter might be ‘More copper is produced when electrodes are closer together,’ or ‘Increasing the amount of product in an electroplating cell can be achieved by decreasing the resistance.’

Using scientific terminology

Every scientific discipline uses different terminology and presents ideas in slightly different ways. In chemistry, you need to convey your ideas by using the correct scientific terminology and aligning them with the way chemical ideas are represented.

Make sure you use correct symbols, equations and units where possible. Personal pronouns are not used when presenting scientific research, so don't include 'I', 'we' or 'us'. Aim to write your poster in third-person past tense.

Referencing

There are many styles of scientific referencing that can be used. Refer to your assessment criteria to see if a particular type is stated. A common referencing style is APA (American Psychological Association), which you can see in Figure 5. You need to include your reference list at the bottom section of your poster.

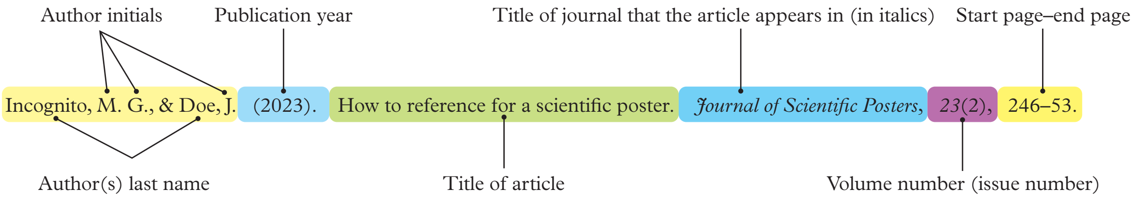


FIGURE 5 An example of how to include a journal article in your reference list using APA referencing

FIGURE 6 An example of how to reference a journal article in the body text of your report

Acknowledgements

The acknowledgements section is where you, the author of the poster, acknowledge and list the people who contributed to or supported your research investigation. People who directly contributed, such as your group members, should appear at the top of the poster as group members and not in the acknowledgements. People you may wish to acknowledge are:

- the lab assistant who prepared all your equipment and materials
- your teacher for contributing ideas
- class members who helped with ideas
- people who helped proofread your poster.

TO-DO LIST

- Write a communication statement to summarise the results of your investigation.
- Use the format of your choice to communicate your findings in 600 words or fewer.
- Format your list of references.
- Include an acknowledgement for everyone who helped you – don't forget your teacher!

CHAPTER 13 Review

Chapter summary

- As part of Unit 4 Area of Study 3, you will conduct a research investigation related to chemical principles explored in Unit 3 and/or Unit 4.
- For your research investigation you will need to collect and analyse primary data; that is, data that has been generated from your own experimentation.
- To answer your investigation question, you can break down the question into smaller parts and generate expanding questions to guide your research.
- Planning your investigation using a research outline is helpful to keeping you on track.
- Annotating your results with detailed notes when collecting data for your investigation can make the research investigation process easier.
- Information can be organised in different ways, but for Unit 4 Area of Study 3, you must present your investigation in VCAA's poster format.
- You must acknowledge your sources by referencing them in the text and in a bibliography.

Research investigation checklist

Use the following checklist to make sure you have completed the research investigation.

I have…	
• written or decided on a scientific question to investigate	☐
• defined the key scientific theories and terms that are relevant to my investigation topic	☐
• analysed the criteria I will be marked on for this assessment	☐
• chosen a methodology for my investigation that will best answer my scientific question	☐
• identified the independent and dependent variables that I will be investigating and identified which variables I will control for my investigation	☐
• written a testable hypothesis	☐
• written a succinct and detailed method for my investigation	☐
• conducted a risk assessment	☐
• evaluated the ethics of my investigation	☐
• set up a logbook to record all results and observations as I conduct the investigation	☐
• determined the best way to display my data to best represent my results	☐
• analysed my results by identifying key findings and linking key findings to theory	☐
• evaluated my errors	☐
• modified the method where needed	☐
• identified if my hypothesis was supported	☐
• identified the limitations of my investigation design and proposed recommendations to overcome limitations	☐
• used a poster format to communicate my findings in 600 words or fewer	☐
• written a communication statement	☐
• formatted my list of references	☐
• included acknowledgements	☐

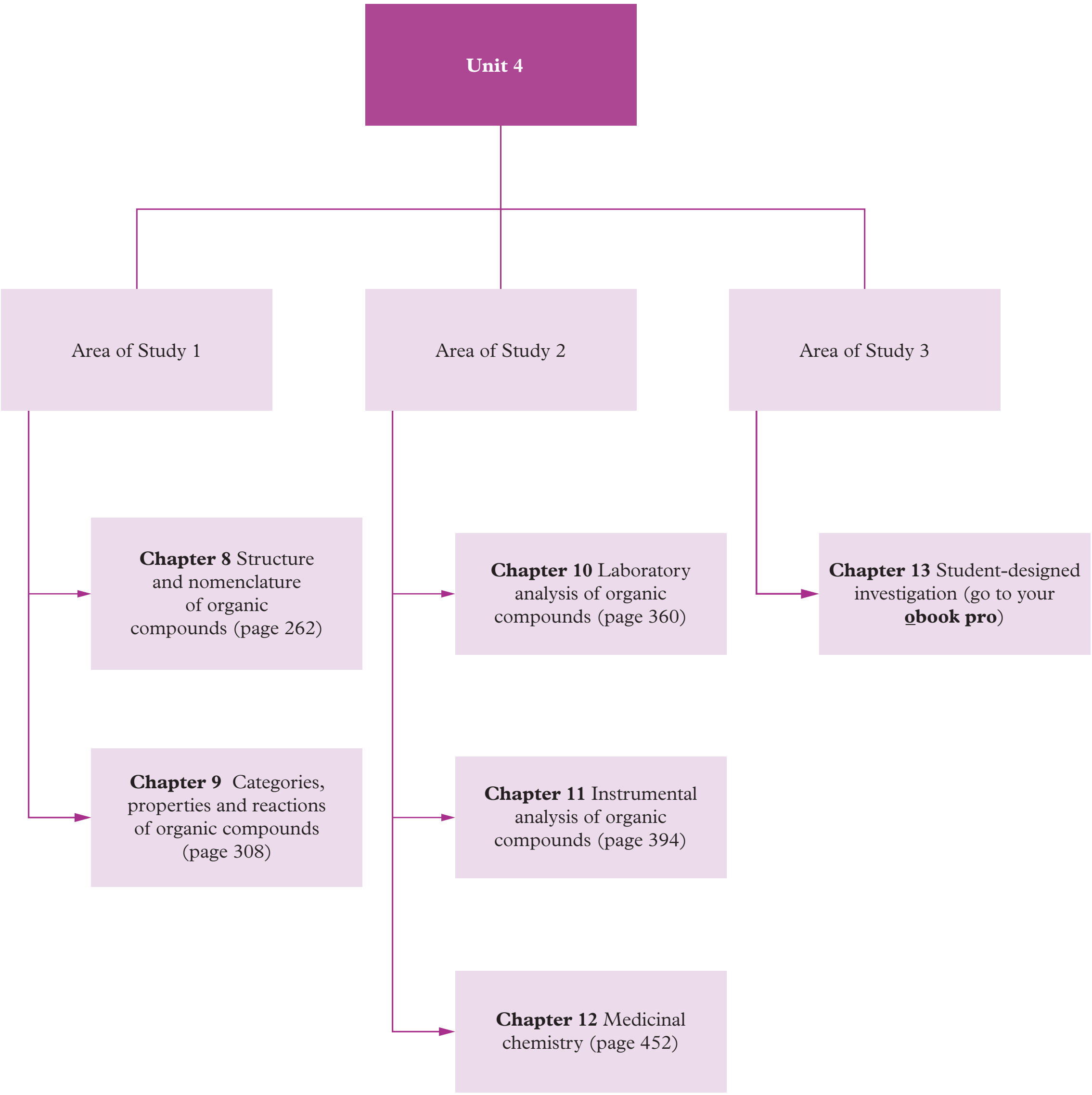
UNIT 4 Review

PART A – Revisit and revise

Part A of the Unit Review will help you revisit and revise all the key concepts and terms from Unit 4 and test your understanding to identify strengths and weaknesses.

Unit 4 overview

The chart below shows all the Areas of Study for Unit 4 and the relevant chapters in your student book. Go to the pages shown to review the key concepts for each chapter.



Test your understanding

Use the following table to guide your revision:

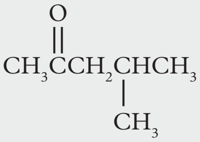
Step 1 – Read the key knowledge for this unit.

Step 2 – Test your understanding of the key knowledge by answering the question(s).

Step 3 – Rate your understanding of each key knowledge from low to high.

Step 4 – Use the topic and page numbers to revise the concepts for which you've identified you need practice.

Only the first three key knowledge dot points are shown. Access the rest of the Test your understanding questions in your **obook pro**.

Key knowledge	Test yourself	Rate yourself	Focus your revision
• characteristics of the carbon atom that contribute to the diversity of organic compounds formed, with reference to valence electron number, relative bond strength, relative stability of carbon bonds with other elements, degree of unsaturation, and the formation of structural isomers	**1** Explain why the element carbon is involved with the formation of many different compounds. Justify your answer with a labelled diagram.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 8.1 Pages 264–270
	2 Identify and determine the degree of unsaturation of five possible isomers of the molecule C_6H_{10}.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 8.1 Pages 264–270
• molecular, structural and semi-structural (condensed) formulas and skeletal structures of alkanes (including cyclohexane), alkenes, benzene, haloalkanes, primary amines, primary amides, alcohols (primary, secondary and tertiary), aldehydes, ketones, carboxylic acids and non-branched esters	**3** Draw the molecular, structural and semi-structural formula for the ester, propyl butanoate.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 8.2 Pages 271–283
• the International Union of Pure and Applied Chemistry (IUPAC) systematic naming of organic compounds up to C8, with no more than two functional groups for a molecule, limited to non-cyclic hydrocarbons, haloalkanes, primary amines, alcohols (primary, secondary and tertiary), aldehydes, ketones, carboxylic acids and non-branched esters	**4** Write the systematic name for the following molecule:  $CH_3\overset{\overset{O}{\|}}{C}CH_2\underset{\underset{CH_3}{\vert}}{CH}CH_3$	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 8.3 Pages 284–289

PART B – Exam essentials

Now that you've completed your revision for Unit 4, it's time to learn and practise some of the skills you'll need to answer exam questions like a pro! To help you, our expert authors have created the following advice and tips to help you maximise your results on the end-of-year examination.

Exam tip 1: Read questions carefully

- In your exam, you may be expected to write balanced equations for reactions or explain the reactions that could be used to identify the functional groups, number of carbon–carbon bonds, or properties of various solutions.
- Make sure that you read the question thoroughly to determine what is required from your answer. For example, observations of qualitative techniques are essential to determine the difference between the types of alcohols. Here, you will need to know the colours of the reactants/products, whether any gases are formed, the pH of the reactant and product solutions, and any reaction conditions.

See it in action

Read the real exam question below and see how the tip has made a difference in the high-scoring and low-scoring responses.

Question 10

> Analytical chemistry deals with methods for determining the chemical composition of samples of matter. A qualitative method yields information about the identity of atomic molecular species of the functional groups in the sample…
>
> Analytical methods are often classified as being either *classical* or *instrumental*.
>
> Source: DA Skoog, FJ Holler and SR Crouch Principles of Instrumental Analysis 6th edition, Thomson Brooks/Cole, Belmont (CA), 2007, p.1

Classical methods include qualitative analysis, such as treating a compound with reagents to observe any reaction, and quantitative methods, such as volumetric analysis, where the amount of a compound is determined by its reaction with a standard reagent.

Instrumental methods include a variety of spectroscopy, such as IR spectroscopy and NMR spectroscopy.

a. Explain how the classical methods of analytical chemistry can be used to determine information about alcohols. In your answer, refer to:

- qualitative analysis and how it can be used to determine whether a compound is an alcohol and, if it is, the type of alcohol
- quantitative analysis.

3 marks

Source: *VCE 2020 Chemistry Exam* reproduced by permission © VCAA

A high-scoring response

a. Explain how the classical methods of analytical chemistry can be used to determine information about alcohols. In your answer, refer to:

- qualitative analysis and how it can be used to determine whether a compound is an alcohol and, if it is, the type of alcohol
- quantitative analysis.

✓

Primary and secondary alcohols are oxidised by acidified solutions of either dichromate or permanganate ($H^+/Cr_2O_7^{2-}$ or H^+/MnO_4^-). Dichromate is yellow and reduces to a green colour when an alcohol is oxidised. Permanganate is purple and reduces to a clear solution when alcohol is oxidised. Tertiary alcohols cannot be oxidised. Therefore, a colour change will occur in primary and secondary alcohols, but not tertiary, which allows you to determine that an alcohol is not tertiary. Once oxidised, the primary alcohol is now a carboxylic acid. This can be determined by reacting it with $NaHCO_3$, which will form CO_2 bubbles.

The concentration of primary or secondary alcohols can be determined by a redox titration with acidified solutions of either dichromate or permanganate.

3 marks

The student has stated how to determine if it is an alcohol, as well as what type of alcohol it is, and is awarded two marks.

The student has addressed the quantitative analysis dot point by using redox titrations as an example.

A low-scoring response

a. Explain how the classical methods of analytical chemistry can be used to determine information about alcohols. In your answer, refer to:

- qualitative analysis and how it can be used to determine whether a compound is an alcohol and, if it is, the type of alcohol
- quantitative analysis.

This answer does not reference colour change or that the permanganate must be acidified. The answer must have a clear link to an observation.

✗

Alcohols are oxidised by permanganate. Only primary and secondary alcohols can be oxidised; tertiary cannot. Alcohols can be titrated.

3 marks

This does not have enough detail regarding what quantitative analysis is and how the titration works/that it is a redox titration.

No explanation is provided on how to determine the difference between a primary and secondary alcohol.

Think like an examiner

To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam. Imagine you are an examiner and use the marking guidance below to mark the response.

Question 1

The amount of iron in a newly developed, heat-resistant aluminium alloy is to be determined.

An 80.50 g sample of alloy is dissolved in concentrated hydrochloric acid and the iron atoms are converted to $Fe^{2+}(aq)$.

This solution is accurately transferred to a 250.0 mL volumetric flask and made up to the mark.

20.00 mL aliquots of this solution are then titrated against a standard 0.0400 M potassium permanganate solution.

$$Fe^{2+}(aq) + MnO_4^-(aq) + 8H^+(aq) \rightarrow Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

Four titrations were carried out and the volumes of potassium permanganate solution used were recorded in the table below.

Titration number	1	2	3	4
Volume of $KMnO_4$	22.03	20.25	21.97	21.99

a. Write a balanced half-equation, including states, for the conversion of MnO_4^- ions in an acidic solution to Mn^{2+} ions.

$MnO_4^-(aq) + 8H^+(aq) + 6e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

2 marks

b. Calculate the average volume, in mL, of the concordant titres of the potassium permanganate solution.

$\frac{20.25 + 21.97 + 21.99}{3} = 21.40$ mL

1 mark

c. Use your answer to part **b** to calculate the amount, in mol, of $MnO_4^-(aq)$ ions used in this titration.

$n(MnO_4^-) = c \times V$

$= 0.0400 \times 21.40$

$= 0.856$ mol

1 mark

d. Calculate the amount, in mol, of $Fe^{2+}(aq)$ ions present in the 250.0 mL volumetric flask.

$n(Fe^{3+})$ in 20 mL aliquot $= n(MnO_4^-) \times 6$

$= 0.856 \times 6$

$= 5.136$ mol

$\frac{250 \text{ mL}}{20 \text{ mL}} = 12.5$

$n(Fe^{2+})$ in 250 mL $= n(Fe^{2+})$ in 20 mL $\times 12.5$

$= 5.136 \times 12.5$

$= 64.2$ mol

2 marks

e. Calculate the percentage, by mass, of iron in the 80.50 g sample of alloy. Express your answer to the correct number of significant figures.

$m(Fe^{2+}) = n \times M$

$= 64.2 \times 55.8$

$= 3582.36\ g$

$\%(m/m) = \frac{3582.36}{80.50} = 44.50\%$

3 marks

Source: *VCE 2010 Chemistry Exam 1* reproduced by permission © VCAA

Marking guide

Question 1a	1 mark for the balanced chemical equation with states
	1 mark for the correct number of electrons
Question 1b	1 mark for the correct calculation with concordant titres
Question 1c	1 mark for using the volume from 1b, converting it to litres and calculating the moles
Question 1d	1 mark for calculating the moles of Fe^{3+} in the 20 mL aliquot
	1 mark for using the dilution factor (250/20) to calculate the moles of Fe^{3+} in 250 mL
Question 1e	1 mark for calculating the mass of Fe
	1 mark for calculating the %*m*/*m* of Fe in the ore by dividing the mass by 80.50
	1 mark for correct answer to the calculation, with significant figures

Note: In chemistry, you can be assigned consequential marks. This means that, if an answer to a previous question is required for another response (i.e. **1b** is needed for **1c**), assessors will complete the calculation using the incorrect value. This ensures that students are not penalised for incorrect values twice.

Fix the response

Consider where you did and did not award marks in the above response. How could the response be improved?

Write your own response to the same question to receive full marks from the examiner.

Exam tip 2: Analyse information, both individually and as a whole, when combining techniques

- In your exam, you may be asked to analyse a set of MS, IR, ^{1}H-NMR and ^{13}C-NMR data, and use this information to identify an unknown organic compound.
- When combining spectroscopy techniques, you must be methodical and look at all the information given, both as a whole and individually – don't let parts of the question lead you to making mistakes further on in the question.

See it in action

Read the real exam question below and see how the tip has made a difference in the high-scoring and low-scoring responses.

Question 8

An unknown organic compound contains carbon, hydrogen and oxygen.

It is known that:

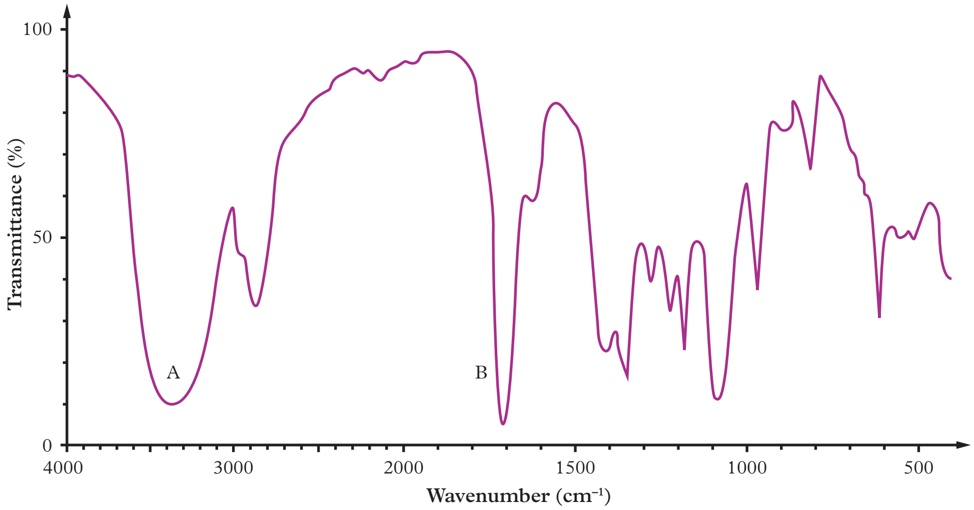
- the compound does not contain carbon–carbon double bonds (C=C)
- the molecular ion peak is found at a mass-to-charge ratio (*m*/*z*) of 74
- the ^{13}C-NMR has three distinct peaks.

b. ii. The ^{1}H-NMR spectrum of the compound shows three sets of peaks with a peak area ratio of 3 : 2 : 1. What does this information tell you about the structure of the compound and its molecular formula? Justify your answer by referring to the information given about the peaks in the ^{1}H-NMR spectrum.

2 marks

c. There are many structural isomers of this compound.
Draw the structural formulas of **two** possible isomers. 2 marks

d. The infrared (IR) spectrum of the compound is shown below.



i. Identify the functional groups responsible for the absorption peak labelled A and B in the IR spectrum.

1 mark

ii. Using the ^{1}H-NMR information in part **b ii** and the IR spectrum provided above, draw the structural formula of the compound. 1 mark

Source: *VCE 2019 Chemistry Exam* reproduced by permission © VCAA

A high-scoring response

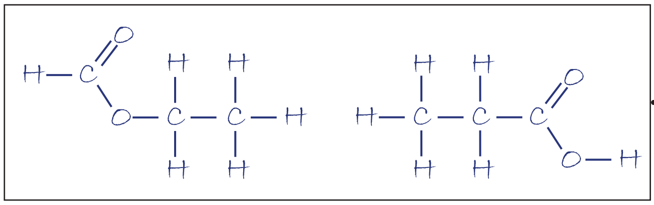
The response has provided two pieces of information that are correct, and is awarded the two marks.

b. ii. There are three hydrogen environments. The ratio tells us the environments contain 3, 2 and 1 hydrogen atoms, so there are six hydrogen atoms in total. Since there are three carbon environments from the ^{13}C-NMR data and no carbon–carbon double bonds, the molecular formula is $C_3H_6O_2$.

2 marks

✓

c.



Both structures show all the bonds for the two isomers structural formulas.

2 marks

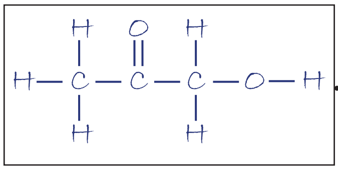
d. i. A: O–H

B: C=O

1 mark

The question is asking for functional groups and not the bonds, so answering *hydroxyl* or *carbonyl* is also acceptable here.

ii.



This was a trick question. On careful inspection, the IR band at A is an alcohol, not a carboxylic acid, so drawing propanoic acid would be incorrect. This is why you need to look at each question individually – don't let previous sections mislead you.

1 mark

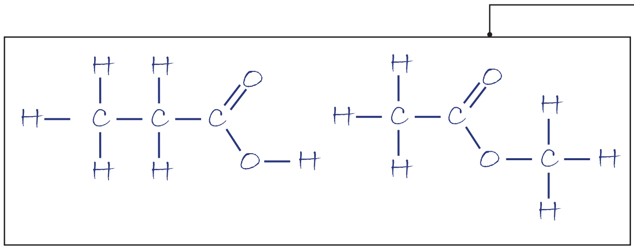
A low-scoring response

This question has two marks for two pieces of information, but the student has only provided one piece of information.

b. ii. Three peaks means there will be three distinct hydrogen environments in this organic molecule.

2 marks

c.



Because the student hasn't looked at the question as a whole, they have missed some integral information and have written the wrong structure. It can't be methyl ethanoate because that only has two hydrogen environments.

2 marks

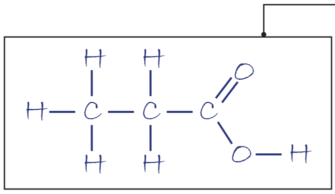
d. i. A: O–H, hydroxyl

B: C=O, carbonyl

Correct responses provided.

1 mark

ii.



Because the student hasn't looked thoroughly at the IR and because they drew propanoic acid in part **c**, they have made an incorrect assumption that the molecule was a carboxylic acid.

1 mark

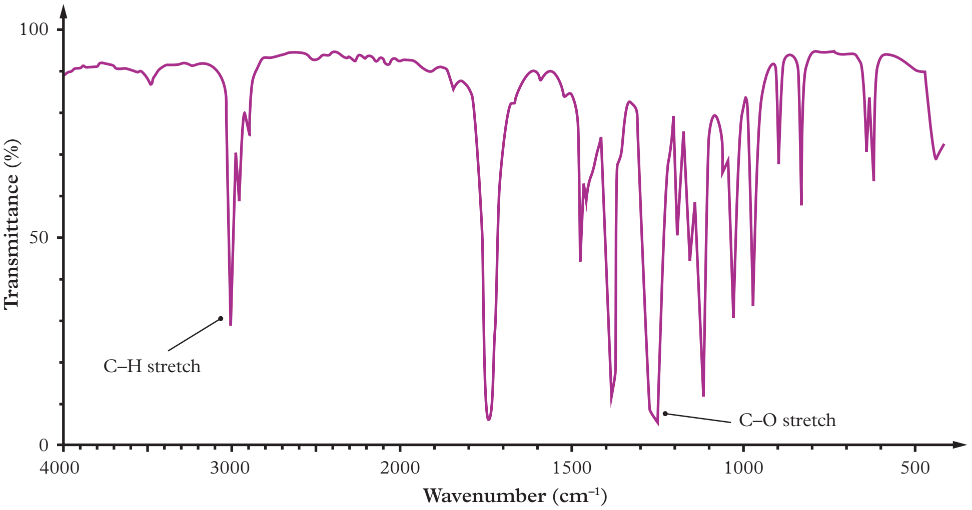
Think like an examiner

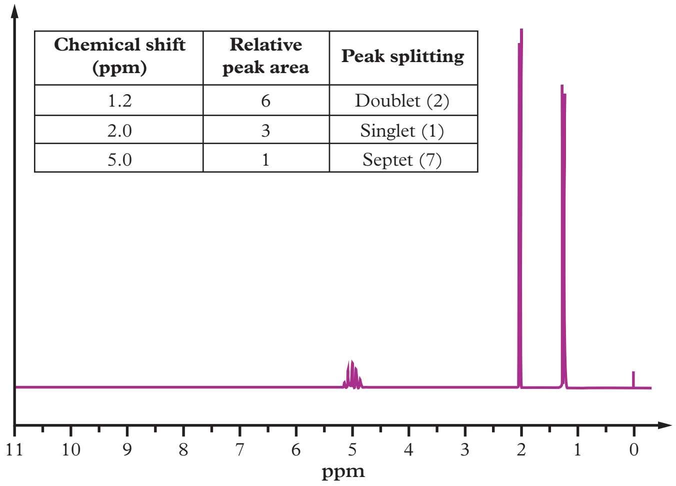
To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam. Imagine you are an examiner and use the marking guidance below to mark the response.

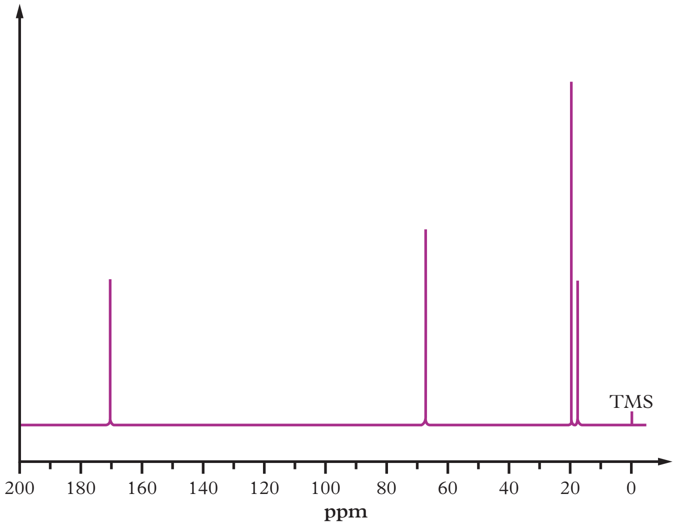
Question 3

While cleaning out a laboratory shelf labelled ‘Carboxylic acids and esters’, a chemist discovers a bottle simply labelled ‘$C_5H_{10}O_2$’. To identify the molecular structure of the contents of the bottle, a sample is submitted for analysis using infrared spectroscopy, and ^{1}H-NMR and ^{13}C-NMR spectroscopy. The spectra are shown below:





Chemical shift (ppm)	Relative peak area	Peak splitting
1.2	6	Doublet (2)
2.0	3	Singlet (1)
5.0	1	Septet (7)



a. Based on the IR spectrum, determine whether the molecule is a carboxylic acid or an ester. Provide a reason for your answer.

The molecule is an ester because of the $C=O$ at 1720–1480.

2 marks

b. Use the information provided in the ^{1}H-NMR and ^{13}C-NMR spectra to identify the number of different chemical environments for hydrogen and carbon in this molecule.

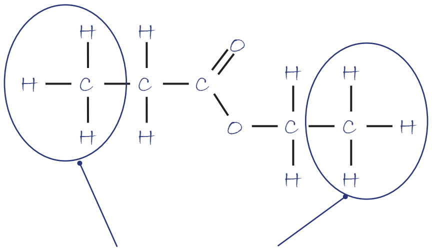
Number of different chemical environments for hydrogen	3
Number of different chemical environments for carbon	4

2 marks

c. Draw the structure of this molecule.

$C_5H_{10}O_2$

- 3H
- 4C
- Ester



Same H and C environments = 3H/4C

2 marks

Source: *VCE 2015 Chemistry Exam* reproduced by permission © VCAA

Marking guide

Question 3a	1 mark for determining that the molecule was an ester
	1 mark for explaining why OH (acid) was not present
Question 3b	1 mark for the correct number of hydrogen environments present
	1 mark for the correct number of carbon environments present
Question 3c	2 marks awarded for the correct structure of the molecule

Fix the response

Consider where you did and did not award marks in the above response. How could the response be improved?

Write your own response to the same question to receive full marks from an examiner.

Practice makes perfect

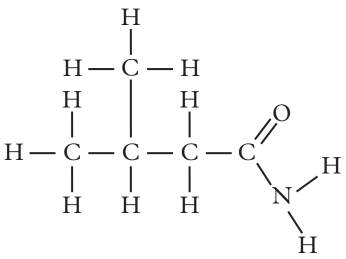
Now that you know all these tips, it's time for you to move on to Part C – Exam practice, to put them into practice.

PART C – Exam practice

Multiple choice

Question 1

What is the name of the compound with the following structural formula?



A 2-methylbutan-1-amide

B 2-methylbutan-4-amide

C 3-methylbutanamide

D 2-methylbutanamide

Question 2

The tertiary structure of proteins may be maintained by:

A hydrogen bonding.

B covalent bonding.

C ionic interactions between NH_3^+ and COO^- groups.

D all of the above.

Question 3

Liquid bromine, Br_2, dissolves in unsaturated hydrocarbons and reacts immediately. Which statement below best describes this process?

A Bromine is polar and reacts by adding bromine atoms across the double bond.

B Bromine is polar and reacts by substituting hydrogen atoms with bromine atoms.

C Bromine is non-polar and reacts by substituting hydrogen atoms with bromine atoms.

D Bromine is non-polar and reacts by adding bromine atoms across the double bond.

Question 4

In response to a pain stimulus, the brain produces small polypeptide molecules called enkephalins. These molecules block the transmission of pain through the central nervous system.

The amino acid sequence in one such compound, methionine enkephalin, is:

Tyr – Gly – Gly – Phe – Met

The number of amine, carboxylic acid and amide (peptide) functional groups in this polypeptide is:

	$-NH_2$	–COOH	–CONH
A	0	0	5
B	1	1	4
C	1	1	5
D	0	2	4

Questions 5 and 6 relate to the following reactions:

Reaction 1: $X + KOH \rightarrow CH_3CH_2OH(l) + KCl$

Reaction 2: $CH_3CH_2OH(l) \xrightarrow{Y} CH_3COOH(l)$

Reaction 3: $CH_3COOH(l) + CH_3OH(l) \rightarrow Z(l) + H_2O(l)$

Question 5

Which of the following shows the formulas of species X, Y and Z?

	Species X	Species Y	Species Z
A	CH_3CH_2Cl	H_2O/H^+	CH_3COCH_3
B	CH_3CH_2Cl	$K_2Cr_2O_7/H^+$	CH_3COOCH_3
C	CH_2ClCH_2Cl	H_2O/H^+	CH_3COOCH_3
D	CH_2ClCH_2Cl	$K_2Cr_2O_7/H^+$	CH_3COCH_3

Question 6

Which of the following correctly identifies each reaction type?

	Reaction 1	Reaction 2	Reaction 3
A	Substitution	Oxidation	Condensation
B	Addition	Oxidation	Reduction
C	Substitution	Elimination	Reduction
D	Addition	Elimination	Condensation

Question 7

The technique of infrared spectroscopy is based on the idea that:

A the nuclei of different atoms are affected by the nuclei of adjoining atoms.

B the bonds between atoms in molecules absorb different wavelengths of energy.

C the bonds between different atoms in a molecule emit different wavelengths of light when excited.

D the bonds between different hydrogen atoms in a molecule absorb different wavelengths of light.

Question 8

Green chemistry is applicable to the ________ of a chemical product.

A life cycle, including design, manufacture and uses

B properties, including design, manufacture and uses

C uses, including design and manufacture

D efficiency, including design, manufacture and uses

Question 9

The reaction that describes the polymerisation of glucose to form cellulose is:

A addition.

B hydrolysis.

C substitution.

D condensation.

Question 10

In green chemistry, atom economy includes:

A using synthetic pathways that incorporate a maximum proportion of the reactant atoms in the product.

B minimising the energy required for a synthetic pathway so that it is more economical.

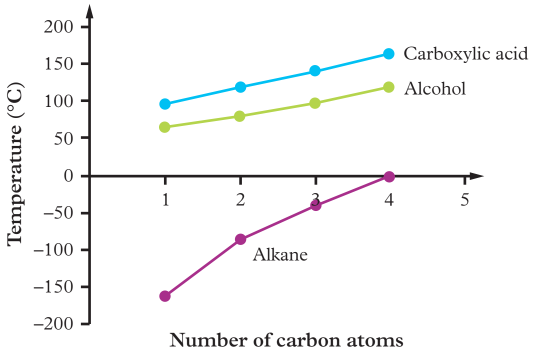
C using catalysts that produce only the desired products and prevent formation of other by-products in a synthetic pathway.

D producing a product that is either biodegradable or recycled at the end of its useful life span.

Short answer

Question 1 (5 marks)

The chart below shows the boiling points of alkanes, alcohols and carboxylic acids with one to four carbon atoms:

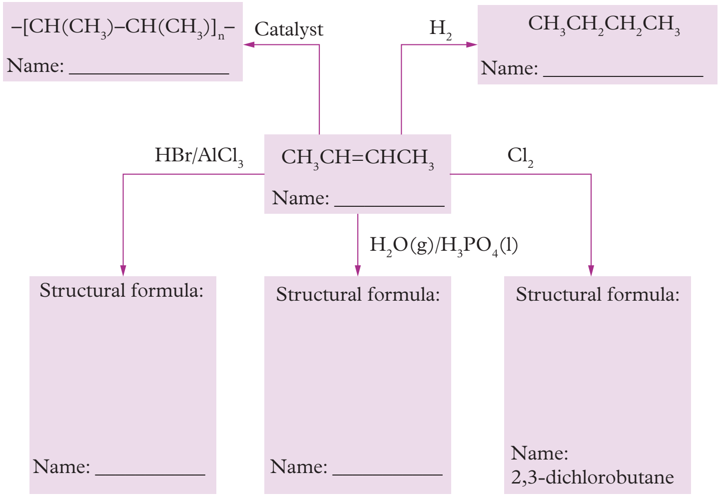


a Explain why the boiling points of the alkanes are much lower than the boiling points of the alcohols and carboxylic acids. (2 marks)

b Explain why the boiling points of all three series of compounds increases with increasing numbers of carbon atoms. (3 marks)

Question 2 (6 marks)

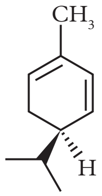
Complete the following flowchart by providing the missing names and structures.



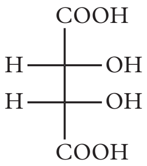
Question 3 (5 marks)

a Deduce the number of chiral carbon atoms in the following molecules.

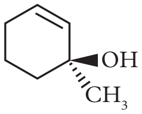
i



ii



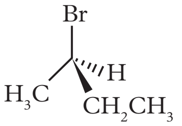
iii



(3 marks)

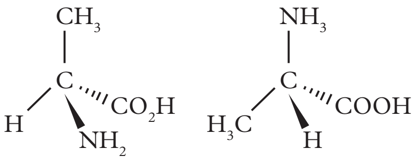
b Deduce whether the following pairs of compounds are stereoisomers or the same compound:

i



(1 mark)

ii



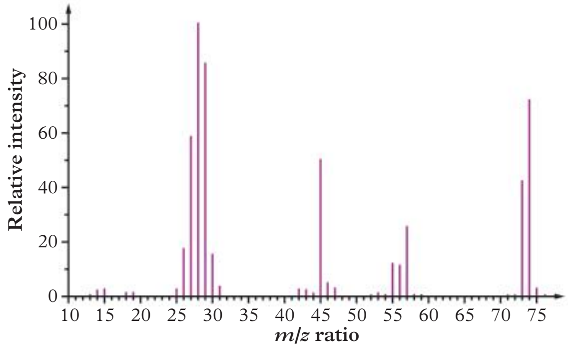
(1 mark)

Question 4 (13 marks)

All data in this question refers to *one compound only*.

a An organic compound was analysed and found to contain 48.6% carbon, 8.1% hydrogen and 43.2% oxygen. Determine the empirical formula of this compound. (3 marks)

b The mass spectrum of the compound in part **a** is given below:



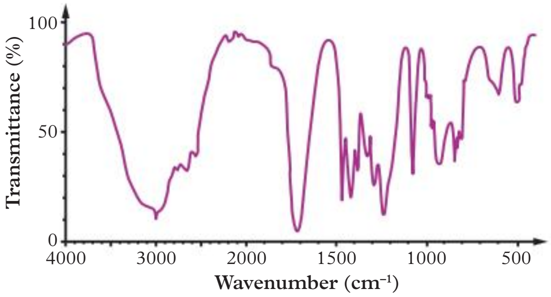
Complete the table below by suggesting a molecular formula for each species that correspond to the *m*/*z* values provided:

***m*/*z* ratio**	29	45	73	74
Species				

(4 marks)

c Deduce the molecular formula of the compound in part **a**. (1 mark)

d The infrared spectrum of the compound in part **a** is given below:



Use the infrared spectrum to deduce the type of chemical bonds in the compound that is responsible for:

i the broad peak at 3050 cm^{-1} (1 mark)

ii the sharp peak at 2920 cm^{-1} (1 mark)

iii the peak at 1720 cm^{-1}. (1 mark)

e Use the information from parts **a**, **b**, **c** and **d** above to determine the structural formula for this compound and apply IUPAC rules to correctly name the compound. (2 marks)

Question 5 (4 marks)

Ethanol is produced in industry by the reaction of ethene with steam according to the following equation and reaction conditions:

$$CH_2CH_2(g) + H_2O(l) \rightleftharpoons CH_3CH_2OH(g) \qquad \Delta H = -45\,kJ\,mol^{-1}$$

$T = 300°C \qquad P = 60\text{–}70\,atm$

Catalyst = phosphoric(V) acid

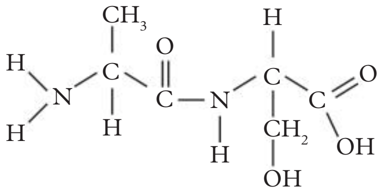
a A compromise is required for choosing the best temperature for this reaction. Deduce a reason for this and propose a resolution. (2 marks)

b Deduce what would occur if the pressure of the reaction were increased. (1 mark)

c Propose reasons why an increase in pressure is not used for this process. (1 mark)

Question 6 (3 marks)

The structure of a dipeptide is shown below:



a Draw a circle on the structure to show the dipeptide link. (1 mark)

b Use your data book to correctly identify the names of the amino acids from which this dipeptide was formed. Redraw the amino acids. (2 marks)

Question 7 (4 marks)

Below are the test results of the miscibility with water for some substances:

Substance	Observations
Hexane (C_6H_{14})	Did not dissolve
Hexene (C_6H_{12})	Did not dissolve
Ethanol (CH_3CH_2OH)	Did dissolve
Propanoic acid (CH_3CH_2COOH)	Did dissolve

Explain these observations with reference to intermolecular forces. (4 marks)

You can find the following resources for this section in your obook pro:

Unit 4 Review Part A
Test your understanding questions

Unit 4 Review Part C
Exam practice questions

Weblink
Past examinations and examiners' reports

Unit 3 & 4 Exam practice
Now that you've completed Units 3 & 4, put your knowledge and skills to the test with these extra practice exam questions.

CHAPTER 14

Practical work

To complete VCE Chemistry, you will need to complete at least 10 hours of practical work for each of Units 3 and 4. Practical work can cover a range of scientific investigation methodologies, such as controlled experiments, modelling, case studies, classification and identification, literature reviews, fieldwork, simulations, and product, process or system development. All investigations that are undertaken as part of your course, as well as school-assessed coursework (SACs), should be written in a logbook that will be monitored and submitted to your teacher. Before undertaking an investigation for the first time, you should consider any ethical concerns, including the importance of sociocultural, economic, political and legal factors that may arise from science-related decision-making.

⚠ SAFETY IN THE LABORATORY

This chapter will highlight key safety concerns for each practical, although there are some general safety concerns to be considered before completing all practical work.

- → Tie long hair back.
- → Do not eat or drink in the lab.
- → Always be aware of your peers and act in a way that will not cause harm.
- → Wear a lab coat, safety glasses, close-toed shoes and gloves.
- → Review the school's safety procedures and the locations of the eyewash, shower, spill kits and first aid kits.
- → Handle all chemicals with care and consult your teacher and risk assessments for all hazards involved with each chemical.
- → Keep open flames away from flammable materials.
- → Handle hot material with the appropriate equipment (i.e. heat-resistant gloves or tongs).
- → Before use, always check that electrical equipment is not damaged and that there are no exposed wires.
- → Conduct fieldwork in groups and complete a full risk assessment before any excursion.

It is the responsibility of the teacher and school to conduct a risk assessment before any practical covered in this book.

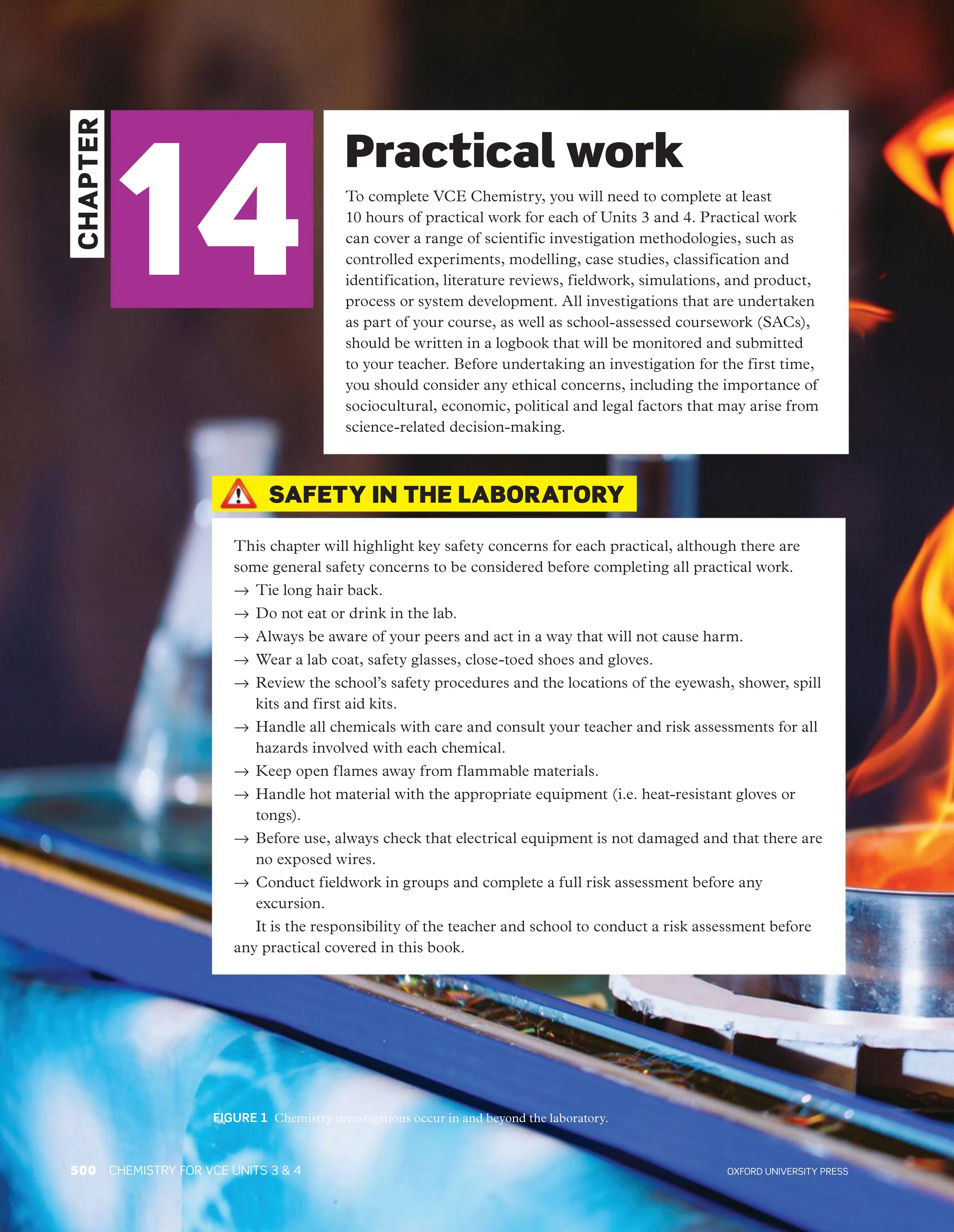
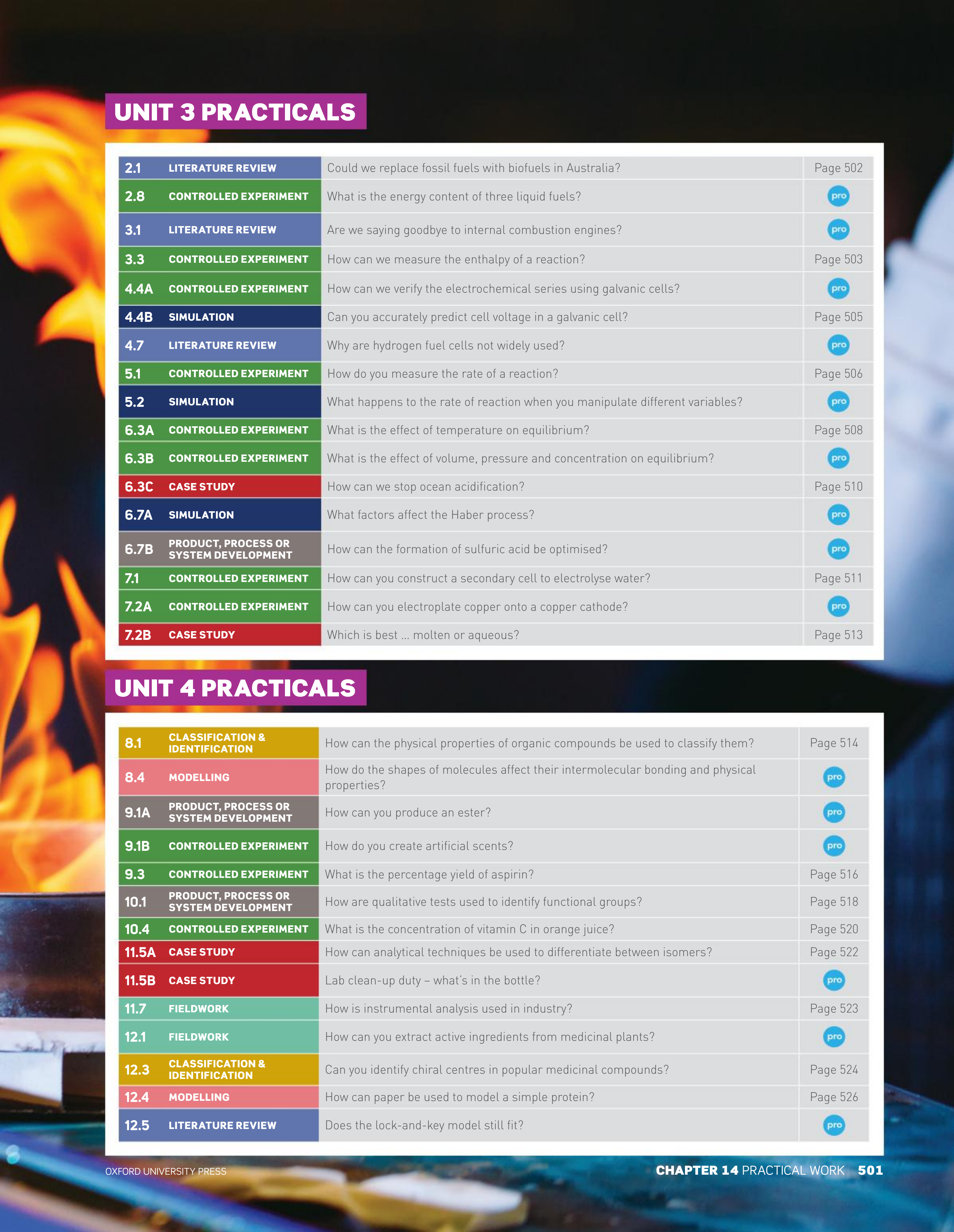


FIGURE 1 Chemistry investigations occur in and beyond the laboratory.



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2.1

PRACTICAL: LITERATURE REVIEW

Could we replace fossil fuels with biofuels in Australia?

Practical worksheet
2.1 Could we replace fossil fuels with biofuels in Australia?

Context

Brazil is a world leader when it comes to production and export of biofuels … but how does Australia compare?

In this practical, you will investigate the benefits of shifting to biofuels as a main source of fuel. You will then use this information to decide whether it is possible for Australia to replace traditional fossil fuels with biofuels.

FIGURE 1 Could we make a complete switch to biofuels in Australia?

Aim

To investigate the production of biofuels in Australia and determine whether it is possible for them to completely replace fossil fuels

Instructions

1 Look for secondary sources of information. These can include the internet, books, scientific magazines, videos, podcasts and interviews with experts.

2 Use the CRAAP method to evaluate the reliability of your sources.

3 Make notes about what you have learnt. You can organise your information in different ways.

4 Prepare a written report that answers the following questions:

  a What are biofuels and how are the different types of biofuels produced?

  b What are the current uses of biofuels? Give statistics where possible.

  c Why are biofuels considered to be renewable?

  d Are biofuels produced in Australia? What can you find out about the demand for biofuels in Australia? How does this compare to the production of biofuel in other countries?

  e Considering all the environmental, social, economic, legal and/or political factors, is it worth producing biofuels in Australia? Is it possible to replace fossil fuels with biofuels? How can we make this possible?

5 Make sure you record the details of all of the sources you use, including the title of the source, who it was written by, when it was written, page numbers or URLs and the date you accessed it.

Refer back to Chapter 13 for some helpful tips on completing your literature review.

2.8

**PRACTICAL:
CONTROLLED EXPERIMENT**

What is the energy content of three liquid fuels?

Video demonstration
2.6 What is the energy content of three liquid fuels?

Practical worksheet
2.6 What is the energy content of three liquid fuels?

Risk assessment
2.6 What is the energy content of three liquid fuels?

Context

Liquid fuels are used widely in internal combustion engines because they are portable, energy dense and flow easily through fuel lines. Liquid fuels can be obtained from crude oil, such as petrol, diesel and kerosene, or from biological sources, such as bioethanol and biodiesel.

In this practical, you will experimentally determine the energy content of a fuel using a technique called calorimetry.

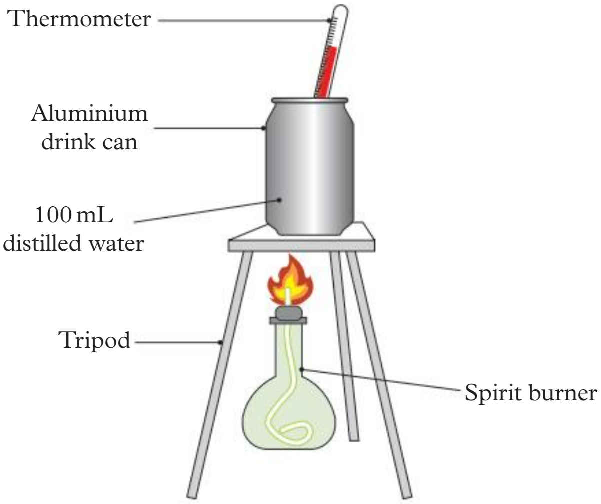


FIGURE 1 Experimental set-up

Aim

To determine the energy content (in kJ g^{-1}) and enthalpy of combustion (in kJ mol^{-1}) of three liquid fuels in a calorimetry experiment.

Materials

- 100 g distilled water
- Aluminium drink can or alternative metal can
- Tripod
- 100°C thermometer
- 1 spirit burner containing methanol
- 1 spirit burner containing ethanol
- 1 spirit burner containing propan-1-ol
- Electronic balance

Method

1. Record the initial mass of the methanol spirit burner with its cap on in Table 1.
2. Set up the experiment according to Figure 1. Weigh out 100 g of water into an aluminium drink can and record its mass. Record the initial temperature of the distilled water in Table 1.
3. Ignite the spirit burner and observe the temperature of the water increase. Extinguish the spirit burner by replacing the cap (do not blow it out) when the temperature of the water has increased by approximately 20°C.
4. Wait for the temperature of the water to stop rising and then record the final temperature of the water in table 1.
5. Record the final mass of the spirit burner with its cap on in Table 1.
6. Repeat the process using new (not warm) distilled water for the other two fuels (ethanol and propan-1-ol).

Results

1. Record your results in Table 1 (see your obook pro). Complete the table to summarise your results.
2. Calculate the heat content of the three fuels in kJ g^{-1} using the formula:

$$\text{heat content of fuel} = \frac{q}{m_{\text{fuel}}}$$

3. For each fuel, calculate the percentage difference between the theoretical enthalpy and the experimental values obtained from your calorimetry experiment.

Discussion

1. Describe any trends you noticed between the ΔH of the fuel and the number of carbon atoms in the fuel.
2. Explain the trends you described in Question 1.
3. Suggest reasons for any differences between theoretical and experimental enthalpy values.
4. Describe three ways to increase the accuracy of the results obtained in this experiment.

Conclusion

Write a suitable conclusion for this experiment.

3.1
PRACTICAL: LITERATURE REVIEW

Are we saying goodbye to internal combustion engines?

Practical worksheet
3.1 Are we saying goodbye to internal combustion engines?

Context

The internal combustion engine was invented in the late 1700s by several scientists, who each unintentionally developed a different piece of the engine. It has long been the best way to generate energy for a variety of machinery.

They have since been re-designed to be smaller, lighter, safer to use, and to start very quickly. Combustion engines have become highly efficient, losing little energy in the form of heat and sound compared to earlier models.

Scientists are now competing to refine newer technologies, including electric cars, which are either battery charged or powered by hydrogen fuel cells. The aim is to create the most cost-effective, fuel efficient and environmentally friendly alternative to the internal combustion engine.

In this practical, you will explore the use of internal combustion engines and compare this to alternative sources of energy.

FIGURE 1 Most vehicles use internal combustion engines

Aim

To investigate internal combustion engines and evaluate whether alternative energy sources are viable replacements

Instructions

1. Look for secondary sources of information. These can include the internet, books, scientific magazines, videos, podcasts, or interviews with experts.
2. Use the CRAAP method to evaluate the reliability of your sources.
3. Make notes about what you have learnt. You can organise your information in different ways.
4. Prepare a 4–5 minute oral presentation video that answers the following questions:
   - **a** What is the main source of fuel for an internal combustion engine?
   - **b** What kind of reaction occurs inside an internal combustion engine? Use a chemical equation and a calculation to show how much carbon dioxide gas can be generated from a full tank of fuel.
   - **c** What are the negative environmental impacts of internal combustion engines? Your answer should refer to renewability, carbon neutrality and pollutants produced.
   - **d** How do battery and hydrogen-based engines work? What reactants and products are involved?
   - **e** Why can battery and hydrogen-based cars be equally bad for the environment as internal combustion engines?
   - **f** Considering all the environmental, social, economic, legal and/or political factors, is it possible to replace internal combustion engines with the more ‘sustainable’ alternatives?
5. Make sure you record the details of all of the sources you use, including the title of the source, who it was written by, when it was written, page numbers or URLs and the date you accessed it.

Refer back to Chapter 13 for some helpful tips on completing your literature review.

3.3
PRACTICAL:
CONTROLLED EXPERIMENT

How can we measure the enthalpy of a reaction?

Video demonstration
3.3 How can we measure the enthalpy of a reaction?

Practical worksheet
3.3 How can we measure the enthalpy of a reaction?

Risk assessment
3.3 How can we measure the enthalpy of a reaction?

Additional resource
Calibrating a calorimeter

Context

Calorimeters provide an effective way of measuring the enthalpy changes in a chemical reaction.

In this practical, you will use calorimetry to measure enthalpy changes.

Before completing Parts A–D of this experiment, your calorimeter will need to be calibrated. This may be completed before or during class. See the Additional resource in your obook pro for the method for calibration.

The calibration factor (CF) is calculated using

$$\mathrm{CF} = \frac{V \times I \times t}{\Delta T}$$

where V is voltage in volts, I is current in amps, t is time in seconds and ΔT is the change in temperature.

Then, the following equations can be used to calculate enthalpy:

$$E = \mathrm{CF} \times \Delta T$$

$$\Delta H = \frac{E}{n}$$

where n is the amount in mol.

Aim

To determine the enthalpy changes (ΔH) of chemical reactions by calorimetry

Materials

- Calorimeter (or alternative shown in Figure 1)
- 50 mL measuring cylinder (or a 50 mL glass Pasteur pipette and pipette bulb to provide more accurate measurements)
- Stopwatch
- Thermometer (digital is faster)
- Electronic balance and weigh boat
- Stirrer
- 150 mL 2.0 M HCl

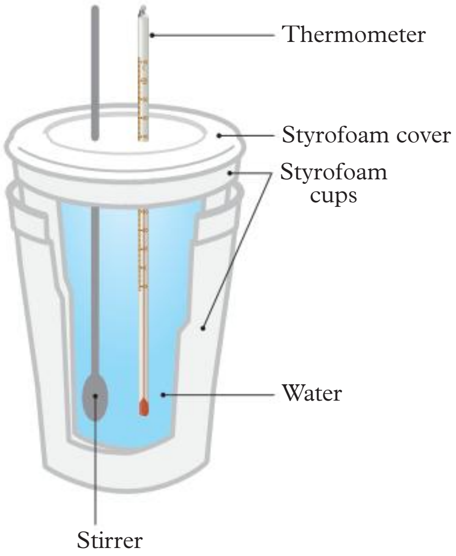


FIGURE 1 An alternative to a calorimeter

- 50 mL 2.0 M NaOH
- 0.9 g MgO
- 3.5 g $MgCO_3$
- 25 mL 1.0 M $CuSO_4$
- 6.0 g zinc powder

Method

Part A – Measure the enthalpy of a neutralisation reaction

1. Add 50 mL of 2.0 M HCl to the calorimeter and record the temperature until it remains stable for 2 minutes. Record this as the initial temperature in Table 1 (see your obook pro).
2. Add 50 mL of 2.0 M NaOH and stir constantly, recording the maximum temperature that the solution reaches in the calorimeter in Table 1.

Theoretical ΔH value is $-58\ \mathrm{kJ\,mol^{-1}}$.

Part B – Measure the enthalpy of a chemical reaction: acid + metal oxide

3 Add 50 mL of 2.0 M HCl to the calorimeter and record the temperature until it remains stable for 2 minutes. Record this as the initial temperature in Table 1.

4 Accurately weigh 0.9 g of MgO, using an electronic balance and weigh boat.

5 Add the MgO to the HCl in the calorimeter and stir constantly, recording the maximum temperature that the solution reaches in the calorimeter in Table 1.

Theoretical ΔH value is $-146\ \text{kJ mol}^{-1}$.

Part C – Measure the enthalpy of a chemical reaction: acid + metal carbonate

6 Add 50 mL of 2.0 M HCl to the calorimeter and record the temperature until it remains stable for 2 minutes. Record this as the initial temperature in Table 1.

7 Accurately weigh 3.5 g of $MgCO_3$ using an electronic balance and weigh boat.

8 Add the $MgCO_3$ to the HCl in the calorimeter and stir constantly, recording the maximum temperature that the solution reaches in the calorimeter in Table 1.

Theoretical ΔH value is $-90\ \text{kJ mol}^{-1}$.

Part D – Measure the enthalpy of a displacement reaction

9 Add 25 mL of 1.0 M $CuSO_4$ to the calorimeter and record the temperature until it remains stable for 2 minutes. Record this as the initial temperature in Table 1.

10 Accurately weigh 6.0 g of Zn powder, using an electronic balance and weigh boat.

11 Add the Zn powder to the $CuSO_4$ in the calorimeter and stir constantly, recording the maximum temperature that the solution reaches in the calorimeter in Table 1.

Theoretical ΔH value is $-217\ \text{kJ mol}^{-1}$.

Results

Record your results in Table 1 (see your obook pro).

Discussion

1 Calculate the CF using the following equation:

$$\text{CF} = \frac{V \times I \times t}{\Delta T}$$

2 Calculate E (in J) for the four reactions and convert it to kilojoules (kJ):

$$E = \text{CF} \times \Delta T$$

3 Calculate the amount, in mol, for each solution by writing a balanced chemical equation for each reaction and determining the limiting reactant.

4 Calculate the change in enthalpy of each of the solutions, using the energy calculated in Question 1 and the amount, in mol, calculated in Question 2.

$$\Delta H = \frac{E}{n}$$

5 Calculate the percentage error for each of the reactions:

$$\%\ \text{error} = \frac{|\text{measured } \Delta H - \text{theoretical } \Delta H|}{\text{theoretical } \Delta H} \times 100\%$$

Note: The |____| symbol tells you to use the absolute value of the number, i.e. remove any negative signs.

6 Explain what errors may be involved with the experiment and describe the specific impact of these errors on the results (i.e. higher/ lower).

7 Describe the assumption being made by using the specific heat capacity of water in the calculations. Determine whether this assumption is justified.

8 Explain why calculating the mass of the solution using its density results in decreased accuracy of the calculations.

Conclusion

Write a suitable conclusion for this experiment.

4.4A

PRACTICAL:
CONTROLLED EXPERIMENT

How can we verify the electrochemical series using galvanic cells?

Video demonstration
4.4A How can we verify the electrochemical series using galvanic cells?

Practical worksheet
4.4A How can we verify the electrochemical series using galvanic cells?

Risk assessment
4.4A How can we verify the electrochemical series using galvanic cells?

Context

Galvanic cells are constructed from two half-cells, a salt bridge, connecting wires and a voltmeter. The reducing and oxidising ability of each half-cell determines the amount of energy produced by each cell and can be measured with the voltmeter. The electrochemical series allows you to predict the outcome of a galvanic cell and the voltage that will result. These half-cells have been measured against the hydrogen half-cell (0.00 V) at 25°C, 1 M and 1 atm.

In this practical, you will construct galvanic cells to experimentally determine the order of metals on the electrochemical series.

Aim

To determine the order of metals on the electrochemical series by constructing galvanic cells made from various metal/metal-ion half-cells

Materials

- 50 mL 1 M $CuSO_4$
- 50 mL 1 M $FeCl_2$
- 50 mL 1 M $Zn(NO_3)_2$
- 50 mL 1 M $Al(NO_3)_3$
- 200 mL 1 M KNO_3
- 4 × 100 mL beakers
- Pen for labelling glassware
- One 2 × 10 cm piece of copper, iron (or a nail), zinc and aluminium metals
- 6 strips of filter paper (approx. 3 × 15 cm)
- 2 × alligator clips
- Plastic tweezers
- Voltmeter
- Emery paper
- Wash bottle with deionised water
- 200 mL waste beaker

Method

FIGURE 1 The experimental set-up of a galvanic cell

1. Construct four half-cells by placing 50 mL of the copper sulfate solution in a 100 mL beaker and adding the copper strip, which acts as the electrode. Repeat this for the iron, zinc and aluminium half-cells. Label each beaker as the respective solution. Remember to scour the metal strips with emery paper.
2. Immerse one piece of the filter paper in the potassium nitrate solution. Using the plastic tweezers, remove it from the solution and place it as a bridge between two of the half-cells (Figure 1).
3. Connect an alligator clip to each of the metal electrodes.
4. Connect the other end of each electrode to the voltmeter. If the voltmeter has a negative reading, swap the wires that are connected to the terminals. **Note:** As soon as the electrodes are connected to the voltmeter, the electrochemical circuit is complete and the voltmeter will immediately measure the voltage of the cell. This must be recorded immediately because it will reduce over time. You must also record whether each electrode is positive or negative in the galvanic cell.

This can be determined by looking at the voltmeter. The negative electrode connects to the negative terminal and the positive electrode connects to the positive terminal.

5. Deconstruct the galvanic cell, ensuring that the filter paper is disposed of in the waste beaker and that no solution in the half-cell contaminates another.
6. Reconstruct the galvanic cell until every pair of half-cells has been connected and the results have been recorded.

Results

Construct a table (such as the one found in your obook pro) that records the voltage produced by each of the six galvanic cells, as well as the polarity of the electrodes.

Discussion

1. Identify the electrode that is always negative and the electrode that is always positive. Use this information to list the half-cells in order, with the strongest oxidising agent first.
2. Compare your half-cell order with the half-equations on the electrochemical series in the VCAA Chemistry data book.
3. Use the electrochemical series to draw the six galvanic cells. You must also add the E^0 of each cell.
4. Compare the theoretical and experimental E^0 values.
5. Identify the solution of the salt bridge. Explain why this ionic solution was selected.
6. On the electrochemical series, copper is the highest half-equation because copper ions are the strongest oxidising agent. Identify the materials that can be used to construct a copper half-cell if it is connected to a tin half-cell. Explain why you chose these materials.

Conclusion

Write a suitable conclusion for this experiment.

4.4B
PRACTICAL: SIMULATION

Can you accurately predict cell voltage in a galvanic cell?

Practical worksheet
4.4B Can you accurately predict cell voltage in a galvanic cell?

Weblink
Teach Chemistry Galvanic cells simulator

Context

The half-cells that make up galvanic cells I–IV are shown below:

I $Ag(s)/Ag^+(aq)$ and $Cu(s)/Cu^{2+}(aq)$
II $Mg(s)/Mg^{2+}(aq)$ and $H_2(g)/H^+(aq)$
III $Zn(s)/Zn^{2+}(aq)$ and $Ag(s)/Ag^+(aq)$
IV $H_2(g)/H^+(aq)$ and $Cu(s)/Cu^{2+}(aq)$

In this practical, you will analyse the galvanic cells and calculate the maximum cell voltage. You will then visualise the electron transfer and ion movement occurring within the galvanic cell, and confirm cell voltage, using a simulation.

FIGURE 1 A voltmeter is attached to a galvanic cell to measure the transfer of electrons.

Aim

To predict cell voltage of various galvanic cells with the help of a simulator

Instructions

1 Draw the galvanic cells in your logbook. Make sure you include:
   - the anode
   - the cathode
   - electrode polarity
   - the materials of the electrodes
   - half-equations for each half-cell
   - overall redox equation
   - direction of electron movement
   - direction of ion movement.

2 Using the electrochemical series, calculate the cell voltage (E^0_{cell}) of galvanic cells I–V:

$$E^0_{cell} = E^0_{reduction} - E^0_{oxidation}$$

3 Use the following link to access the Teach Chemistry: Galvanic cells simulator: https://teachchemistry.org/classroom-resources/voltaic-cells

4 Determine which half-cell should be the Left Beaker and which should be the Right Beaker. The Left Beaker contains the positive electrode (cathode) and is connected to the positive terminal. The Right Beaker contains the negative electrode (anode) and is connected to the negative terminal.

5 Use the drop-down menus to select the electrode material and electrolyte solution for the Left Beaker and Right Beaker to create galvanic cells I–IV.

6 Turn on the voltmeter to see the reaction begin. Record the cell voltages for each galvanic cell.

7 Click the ‘See Molecular Scale’ buttons to view electron transfer and ion movement in the Left Beaker, Right Beaker and Salt Bridge.

Questions

1 For each galvanic cell, describe what is happening in the:
   a Left Beaker
   b Right Beaker
   c Salt Bridge.

2 Explain why the hydrogen electrode is in hydrochloric acid, HCl.

3 Explain why the hydrogen half-cell uses an inert electrode.

4 Identify one limitation of the simulation.

4.7
PRACTICAL: LITERATURE REVIEW

Why are hydrogen fuel cells not widely used?

Practical worksheet
4.7 Why are hydrogen fuel cells not widely used?

Context

Despite the obvious environmental benefits of hydrogen fuel cells, they are not widely used.

In this practical, you will explore the (limited) uses of hydrogen fuel cells and investigate alternatives that are more sustainable than fossil fuels.

FIGURE 1 Hydrogen fuel station

Aim

To evaluate hydrogen fuel cells and the alternative sources of energy and determine which is the most feasible to help society meet its energy needs

Instructions

1 Look for secondary sources of information. These can include the internet, books, scientific magazines, videos, podcasts, or interviews with experts.

2 Use the CRAAP method to evaluate the reliability of your sources.

3 Make notes about what you have learnt. You can organise your information in different ways.

4 Prepare a written report that answers the following questions:
  - **a** How do hydrogen fuel cells work?
  - **b** What are the advantages of using hydrogen fuel cells over fossil fuels derived from crude oil?
  - **c** What factors limit the widespread use of hydrogen fuel cells?
  - **d** What sustainable alternatives are there to hydrogen fuel cells?
  - **e** Considering all the safety, environmental, social, economic, legal and/or political factors, which source of fuel is the most desirable?

5 Make sure you record the details of all the sources you use, including the title of the source, who it was written by, when it was written, page numbers or URLs and the date you accessed it.

Refer back to Chapter 13 for some helpful tips on completing your literature review.

5.1
PRACTICAL: CONTROLLED EXPERIMENT

How do you measure the rate of a reaction?

Video demonstration
5.1 How do you measure the rate of a reaction?

Practical worksheet
5.1 How do you measure the rate of a reaction?

Risk assessment
5.1 How do you measure the rate of a reaction?

Context

You can measure how changing different factors affects the rate of a chemical reaction.

In this practical, you will investigate the effects of surface area, temperature or concentration on the rate of reaction between calcium carbonate, $CaCO_3$, and hydrochloric acid, HCl.

$CaCO_3$ and HCl react to form a solution of calcium chloride, $CaCl_2$, and carbon dioxide gas, CO_2, according to the following chemical equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

According to the law of conservation of mass, the mass lost by the calcium carbonate should be equal to the mass of the carbon dioxide generated.

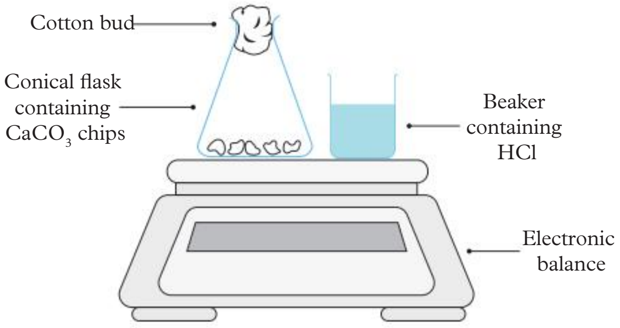


FIGURE 1 The experimental set-up

Aim

To determine the factors which increase the rate of a chemical reaction

Materials

- 5.00 g small $CaCO_3$ chips
- 20 mL of 1 M HCl
- 100 mL conical flask
- 100 mL beaker
- Cotton wool
- Electronic balance
- Spatula
- 20 mL measuring cylinder
- Stopwatch
- Thermometer

For inquiry

- 5.00 g large $CaCO_3$ chips
- 5.00 g $CaCO_3$ powder
- 20 mL of 2 M HCl
- 20 mL of 0.5 M HCl
- Heating plate or hot water bath
- Ice bath
- Heat-proof gloves or alternative to avoid burns

Method

1 Place the conical flask on the electronic balance and tare the balance. Using the spatula, measure approximately 5.00 g of the chips into the conical flask. (Note: You may not be able to measure exactly 5.00 g, but aim to get as close as possible). Record the mass of the chips in your logbook.

2 Measure 20 mL of 1 M HCl using the measuring cylinder. Ensure that the bottom of the meniscus is on the 20 mL line at eye level. Pour the 20 mL of HCl into the beaker. Measure the temperature of the acid using the thermometer and record this value in your logbook.

3 Place both the conical flask and the beaker on the balance. Place the cotton bud on top of the mouth of the conical flask and tare the balance so that it reads 0.00 g.

4 Prepare a stopwatch to start timing in this step. You will record the mass on the balance every 15 seconds.

Lift the cotton bud, then add the acid into the conical flask and start timing immediately. Place the beaker back on the balance and the cotton bud in the mouth of the conical flask.

5 Record the mass of the balance every 15 seconds until the mass remains unchanged for 60 seconds.

6 Dispose of chemicals appropriately, as per your teacher's instructions. Make sure that no chemicals are disposed of down the sink and that the acid and the $CaCO_3$ chips do not enter the drain.

7 Select one variable to modify (see Inquiry below).

Results

1 Record the mass lost every 15 seconds using a results table in your logbook. Remember that the mass at time zero was 0.00 g.

2 Create a line graph of the mass of gas lost every 15 seconds in your logbook.

Discussion

1 Calculate the amount, in mol, of $CaCO_3$ based on the mass of the chips that you weighed.

2 Calculate the amount, in mol, of 1 M HCl based on the volume of the acid used.

3 Identify the reactant that is in excess in the reaction.

4 Calculate the theoretical mass of CO_2 that should be generated by the experiment.

5 Compare the theoretical and experimental masses of CO_2 generated in the experiment and account for any differences.

6 Explain how the trend represented on the graph represents the rate of a chemical reaction.

7 Explain why the rate of the reaction eventually becomes zero.

Inquiry: Choose one of the following questions to investigate.

- What if the concentration was increased or decreased?
- What if the temperature was changed?
- What if the surface area was changed?

1 Write a hypothesis for your investigation.

2 Identify the independent variable.

3 Identify the dependent variable.

4 Identify two key controlled variables.

5 Write a list of materials that you will need in your experiment.

6 Write a method that you will follow in your experiment. This can be amended from the experimental method above.

7 Prepare a table to record your results.

8 Graph each reaction on the same set of axes.

9 For each graph, explain the impact that either surface area, concentration or temperature has on the rate of a reaction.

10 Explain whether temperature can be controlled in the experiment. Describe what impact this could have on the results. Identify one way that impacts could be minimised.

11 Identify two errors and explain the impact that they have on the resulting data. Describe how they can be improved in future experiments.

12 Evaluate the accuracy, precision and validity of the experiment. Justify your responses. Describe how each could be improved.

Conclusion

Write a suitable conclusion for this experiment.

5.2
PRACTICAL: SIMULATION

What happens to the rate of reaction when you manipulate different variables?

Practical worksheet
5.2 What happens to the rate of reaction when you manipulate different variables?

Weblink
Teach Chemistry Reaction Rates simulator

Context

The rate of reaction is affected by a range of variables, including:

- temperature
- surface area
- concentration
- catalysts.

In this practical, you will use a simulation to explore the effects of these variables on the reaction rate.

FIGURE 1 Porous nickel catalysts are commonly used in chemical processes.

Aim

To predict the impact of manipulating variables on the reaction rate

Instructions

Getting started

1 Use the following link to access the Teach Chemistry Reaction Rates simulator: https://teachchemistry.org/classroom-resources/reaction-rates-simulation

2 Use the 'RUN REACTION' button to see how the A+B→C reaction proceeds.

3 You can explore the effects of changing different variables using the 'SELECT VARIABLES →' button. But, before you do, work through the questions below.

Questions

1 The reaction curve is plotted as amount of product vs time. Identify the part of the graph that shows the reaction rate.

2 If one coloured dot represents one mole of reactant, determine the amount, in mol, of product formed in this particular reaction.

3 Explain whether you would expect manipulating the variables above to affect the yield of the product. Justify your response.

Part A – Temperature

4 Redraw the CONTROL reaction curve in your logbook. Predict the impact of the following changes on the rate of reaction, by drawing curves on the same graph:
  - **a** decreasing temperature
  - **b** increasing temperature.

5 Provide a justification for your predictions in Question 4.

6 Identify whether your predictions were correct.

Part B – Surface area

7 Redraw the CONTROL reaction curve in your logbook. Predict the impact of the following changes on the rate of reaction, by drawing curves on the same graph:
   a decreasing surface area
   b increasing surface area.

8 Provide a justification for your predictions in Question 7.

9 Identify whether your predictions were correct.

Part C – Concentration

10 Redraw the CONTROL reaction curve in your logbook. Predict the impact of the following changes on the rate of reaction, by drawing curves on the same graph:
   a decreasing concentration
   b increasing concentration.

11 Provide a justification for your predictions in Question 10.

12 Identify whether your predictions were correct.

Part D – Catalyst

13 Redraw the CONTROL reaction curve in your logbook. Predict the impact of adding a catalyst on the rate of reaction, by drawing a curve on the same graph.

14 Provide a justification for your prediction in Question 13.

15 Identify whether your predictions were correct.

6.3A PRACTICAL: CONTROLLED EXPERIMENT

What is the effect of temperature on equilibrium?

Video demonstration
6.3A What is the effect of temperature on equilibrium?

Practical worksheet
6.3A What is the effect of temperature on equilibrium?

Risk assessment
6.3A What is the effect of temperature on equilibrium?

Context

In this practical, you will analyse three equilibrium systems:

(1) $Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons FeSCN^{2+}(aq)$ ΔH = exothermic

Iron(III) is yellow and iron thiocyanate is dark red.

(2) $H_3PO_4(aq) \rightleftharpoons H_2PO_4^{-}(aq) + H^{+}(aq)$ ΔH = exothermic

Using methyl violet indicator, an increase in hydronium concentration results in a yellow colour. A lower hydronium concentration results in a green/blue colour.

(3) $\underset{\text{brown}}{I_2(aq)} + \underset{\text{clear}}{\text{starch}} \rightleftharpoons \underset{\text{blue–black}}{\text{starch–iodine complex(aq)}}$ ΔH = exothermic

Iodine is brown, whereas starch is clear and the starch–iodine complex is dark blue–black.

(4) $\underset{\text{colourless}}{N_2O_4(g)} \rightleftharpoons \underset{\text{dark brown}}{2NO_2(g)}$ ΔH = endothermic

FIGURE 1 The solution turns blue–black in the presence of starch indicator and I_2.

Depending on whether a reaction is endothermic or exothermic, the position of equilibrium will shift differently when there is a change in temperature (heat energy).

Aim

To determine the effect of changing the temperature in exothermic and endothermic reaction systems on the position of equilibrium

Materials

- 30 mL of 0.005 M $Fe(SCN)^{2+}$ solution (a mixture of $Fe(NO_3)_3$ and KSCN) in a dropper bottle
- 30 mL of 1 M phosphoric acid (H_3PO_4)
- A methyl violet indicator
- 20 mL starch solution
- A small dropper bottle of iodine solution (0.005 M)
- Three stoppered test tubes filled with a mixture of NO_2/N_2O_4 gases
- Nine small or semi-micro test tubes
- One 100 mL beaker
- Three 250 mL beakers, one half-filled with ice water and one half-filled with hot water from a kettle. **Note:** The beakers must hold the test tubes. Ensure the water level will not cause the test tubes to float.
- Stirring rod
- Kettle or a hot plate
- Ice
- Test tube rack

Method

1 Prepare the three 250 mL beakers so that one is empty (this will be the room temperature beaker), one is half-filled with iced water and one is half-filled with hot water.

2 Add the $Fe(SCN)^{2+}$ solution to three of the test tubes until they are half-full and place them in the test tube rack.

3 Add two drops of methyl violet to three test tubes. Then, add the H_3PO_4 solution until they are half-full and place them in the test tube rack. **Note:** If the methyl violet is placed on top of the H_3PO_4 solution, it can be hard to mix. Add the methyl violet first.

4 Measure 20 mL of starch solution into the 100 mL beaker and add 2–3 drops of iodine, then stir until the solution is dark blue. Pour the solution into three test tubes until they are half-full and place them in the test tube rack.

5 Place one test tube containing each solution in the ice bath beaker, the room temperature beaker and the hot water bath (see Table 1).

TABLE 1 Three test tubes in each beaker, each containing different solutions

Ice bath beaker	Room temperature beaker	Hot water beaker
$Fe(SCN)^{2+}$	$Fe(SCN)^{2+}$	$Fe(SCN)^{2+}$
H_3PO_4	H_3PO_4	H_3PO_4
Starch/iodine	Starch/iodine	Starch/iodine

6 Record the colours of each solution at room temperature, in the low temperature ice bath and the hot temperature beaker in Table 2 (see your obook pro).

Note: Steps 7 and 8 can be completed as a demonstration and set up in advance in the fume hood so that students can observe the results and record them in their logbooks without needing to access the fume hood.

7 Remove the test tubes from the beakers and place them in the test tube rack. Take the three beakers to the fume hood where the stoppered test tubes filled with a mixture of NO_2/N_2O_4 gases must be tested. The NO_2/N_2O_4 gases MUST NOT be removed from the fume hood.

8 Place one of each of the test tubes in each beaker in the fume hood and leave them for a couple of minutes. Record the changes in Table 2.

Results

Record your results in Table 2 (see your obook pro).

Discussion

1 For each test, use the results to explain whether the addition of heat will cause a net forward or net reverse reaction.

2 For each test, use the results to explain whether the removal of heat will cause a net forward or net reverse reaction.

3 Use Le Chatelier's principle to explain the colour changes that occur in each test tube. Your answer must refer to exothermic/endothermic reactions, temperature, a partial opposition to a change, and a net forward or reverse reaction.

4 Draw a graph for each hot and cold test tube which represents the initial equilibrium, the change which occurs, and the new equilibrium. Note: Each graph will illustrate the general change because no concentration numbers can be determined.

Inquiry: What if temperature was increased or decreased?

1 Identify the independent variable that you will need to change to test your hypothesis.

2 Identify the dependent variable that you will need to measure in your investigation.

3 Write a hypothesis for your investigation.

4 Identify two controlled variables that will affect the dependent variable unless they are kept constant.

5 Explain the expected results using Le Chatelier's principle.

Conclusion

Write a suitable conclusion for this experiment.

6.3B

PRACTICAL: CONTROLLED EXPERIMENT

What is the effect of volume, pressure and concentration on equilibrium?

Video demonstration
6.3B What is the effect of volume, pressure and concentration on equilibrium?

Practical worksheet
6.3B What is the effect of volume, pressure and concentration on equilibrium?

Risk assessment
6.3B What is the effect of volume, pressure and concentration on equilibrium?

Context

In this practical, you will explore two different equilibrium systems to investigate the effect of volume, pressure and concentration on the position of the equilibrium reactions.

To investigate the effect of volume and pressure on equilibrium, the following system will be studied:

$$\underset{\text{colourless}}{N_2O_4(g)} \rightleftharpoons \underset{\text{dark brown}}{2NO_2(g)} \qquad \Delta H = \text{endothermic}$$

To investigate the effect of concentration on equilibrium, the following system will be studied:

$$\underset{\text{yellow}}{Fe^{3+}(aq)} + \underset{\text{colourless}}{SCN^-(aq)} \rightleftharpoons \underset{\text{red}}{FeSCN^{2+}(aq)} \qquad \Delta H = \text{exothermic}$$

The initial equilibrium mixture should be orange in colour because it is a mixture of the yellow and red colours. As reactants or products are added or removed, the system will partially oppose the change and move to one side of the reaction in a net forward or reverse direction. A net forward reaction will cause the system to become more red or darker, and a net reverse reaction will cause the system to become more yellow or lighter.

Aim

To determine the effect of changing volume, pressure and concentration on the position of equilibrium.

Materials

- NO_2/N_2O_4 gas mixture
- Large gas syringe
- Rubber stopper for the nozzle of the syringe (ensure it is properly sealed and no gas can escape)
- 30 mL 0.005 M $FeSCN^{2+}$ solution (a mixture of $Fe(NO_3)_3$ and KSCN in a dropper bottle)
- 0.1 M $Fe(NO_3)_3$ in a dropper bottle
- 0.1 M KSCN in a dropper bottle
- 0.1 M NaF in a dropper bottle
- 0.1 M $AgNO_3$ in a dropper bottle
- Wash bottle filled with deionised water
- 5 micro test tubes
- Micro test-tube rack

Method

This practical is to be completed as a teacher demonstration.

Part A – The effect of volume and pressure

1 Ensure that the plunger is situated so that it can be both pulled out and pushed in without any gas escaping.

2 While in the fume cupboard, hold the syringe and ask students to record the colour.

3 While holding the sealed end of the syringe (to ensure there are no leaks), press the plunger inwards to compress the gas and hold it in for at least 5 seconds. Students must record the initial colour change (when the plunger went in) and the final colour change that happens after that.

4 Again holding the sealed end of the syringe (to ensure there are no leaks), pull the plunger outwards to reduce the pressure of the gas and hold it there for at least 5 seconds. Students must record the initial colour change (when the plunger was pulled out) and the final colour change that happens after that.

Part B – The effect of concentration

5 Label the five test tubes as A, B, C, D and E and place them in the test-tube rack.

6 Add the $FeSCN^{2+}$ solution to all test tubes so that they are one-third full. Record the initial colour of the solution in the test tubes in Table 1 (see your obook pro).

7 Add solutions to each test tube as follows:

A: Add one drop of $Fe(NO_3)_3$.

B: Add one drop of KSCN.

C: Add one drop of NaF.

D: Add one drop of $AgNO_3$.

E: Add an equal volume of deionised water (total volume of the test tube should be two-thirds full).

8 Record the colour changes of each test tube in Table 1. Note: If the Fe^{3+} is pale yellow and the $FeSCN^{2+}$ is red, the colour changes should be lighter/yellow, darker/red or no change.

Results

Part A – The effect of volume and pressure

1 Record the initial colour of the gas mixture.

2 Record your results in Table 1. Use the terms ‘lighter’ or ‘darker’ to compare the final colours with the initial colour.

Part B – The effect of concentration

3 Record your results in Table 2 (see your obook pro).

Discussion

Part A – The effect of volume and pressure

1 Explain why two colour changes occur when the volume of the system is changed.

2 Explain the colour changes that occur when there is an increase in volume, by referencing Le Chatelier’s principle.

3 Explain the colour changes that occur when there is a decrease in volume, by referencing Le Chatelier’s principle.

4 In a separate reaction, colourless hydrogen gas and dark-purple iodine gas are reacted to form colourless hydrogen iodide gas, according to the following equation:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

a Explain the colour changes that occur when there is an increase in volume, by referencing Le Chatelier’s principle.

b Explain the colour changes that occur when there is a decrease in volume , by referencing Le Chatelier’s principle.

Part B – The effect of concentration

5 For each test, state how the addition of each substance changes the concentration of the reaction system (e.g. lowers the concentration of $FeSCN^{2+}$) and whether a net forward or reverse reaction results, based on the data you have gathered.

6 Using the short-answer structure, use Le Chatelier’s principle to explain the colour changes that occur in each test tube. Your answer must reference exothermic/endothermic, energy, a partial opposition to a change, and a net forward or reverse reaction to release or store energy from the bonds of the reactants or products.

7 Draw a graph for each test tube which represents the initial equilibrium, the change that occurs and the new equilibrium. Note: This graph will illustrate the general change because no concentration numbers are able to be determined.

8 State the equilibrium expression for the reaction.

Conclusion

Write a suitable conclusion for this experiment.

6.3C
PRACTICAL: CASE STUDY

How can we stop ocean acidification?

Practical worksheet
6.3C How can we stop ocean acidification?

Context

Atmospheric carbon dioxide (CO_2) dissolves readily in water. In aqueous form, the carbon dioxide will react with water to form carbonic acid (H_2CO_3):

$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

When the atmospheric concentration of CO_2 increases, more CO_2 will become dissolved in water. Carbonic acid rapidly decomposes into the bicarbonate ion (HCO_3^-) and hydrogen ion (H^+).

$$H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$$

A further reaction occurs when the bicarbonate ion dissociates into the carbonate ion (CO_3^{2-}) and the hydrogen ion (H^+).

$$HCO_3^-(aq) \rightleftharpoons CO_3^{2-}(aq) + H^+(aq)$$

Minor changes in the pH of the ocean can cause drastic changes to marine and aquatic life. As the carbonate ion (CO_3^{2-}) is released into the oceans, it binds to calcium ions (Ca^{2+}), which aquatic and marine life use to build and strengthen their shells.

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightleftharpoons CaCO_3(s)$$

Because the CO_3^{2-} is being used to form $CaCO_3$, there are fewer calcium ions available to build coral and shells.

FIGURE 1 A mollusc shell dissolves under acidic conditions.

Aim

To assess the impact of increased atmospheric CO_2 on the pH of the ocean and propose methods to stop ocean acidification

Questions

1 List five ways in which humans are contributing to the increase in atmospheric CO_2.

Refer to Le Chatelier's principle in your responses to Questions 2–4.

2 Describe the impact of increased CO_2 concentration on the concentration of H_2CO_3 in the ocean.

3 Describe the impact of increased concentration of H_2CO_3 in the ocean on its pH.

4 Describe the impact of increased HCO_3^- on the pH of the ocean.

5 Use your answers to Questions 2–4 to describe how the increase in CO_2 emissions affects the pH of the ocean.

6 Research the changes in pH levels of the ocean over the last 50 years. Explain whether you consider this change to be significant and suggest a reason why or why not.

7 Explain why the weakening of shells of aquatic life cannot be reversed.

8 Discuss and evaluate recommendations you could make to the Australian government to stop ocean acidification and return the ocean's pH to pre-1800 levels (~8.2).

9 Discuss the sustainable development goals that this practical relates to.

6.7A

PRACTICAL: SIMULATION

What factors affect the Haber process?

Practical worksheet
6.7A What factors affect the Haber process?

Weblink
Chemical equilibrium in the Haber process by Wolfram simulation

Context

The Haber process is used to generate ammonia to fertilise crops for the mass production of food. It is a reaction between hydrogen gas and nitrogen gas to produce ammonia.

$$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92\,\text{kJ}\,\text{mol}^{-1}$$

The reaction conditions can be modified to optimise the yield of ammonia.

In this practical, you will use a simulation to explore how different factors affect the equilibrium concentrations of substances involved in the Haber process.

FIGURE 1 Ammonia is used to fertilise crops.

Aim

To investigate the effect of pressure, temperature and concentration on the equilibrium concentrations of the Haber process reaction mixture

Instructions

Getting started

1 Use the following link to access the Chemical equilibrium in the Haber Process by Wolfram simulation: http://demonstrations.wolfram.com/ChemicalEquilibriumInTheHaberProcess/

2 When using the sliders to adjust values, you may need to wait for several seconds until the equilibrium system responds.

Part A – Effect of temperature

3 Set the pressure to 200 bar (20 000 kPa) and the temperature to 600 K by using the sliders at the top. Add 1 mol of each substance by using the sliders at the top. Use the sliders to increase and decrease the temperature, and record the equilibrium amounts in Table 1 (see your obook pro). Nitrogen has already been added to the table as 0.82 mol.

Part B – Effect of pressure

4 Leave the pressure at 200 bar (20 000 kPa), and the amounts of the three reactants at 1 mol each. Reset the temperature to 600 K. The amount of nitrogen at equilibrium is already in Table 2 (see your obook pro). Use the sliders to increase the pressure to 250 bar (25 000 kPa) and record the equilibrium amounts of the three components (in mol) in Table 2. Reduce the pressure and record the amounts of each species.

Part C – Effect of concentration

5 Reset the pressure to 200 bar (20 000 kPa), and the amounts of the three reactants to 1 mol each. Ensure the temperature is still at 600 K. Use sliders to add additional moles of nitrogen, hydrogen and ammonia at constant pressure of 250 bar and constant temperature (600 K), and observe how they change the position of equilibrium, recording your results in tables 3–5 (see your obook pro).

Questions

1 Deduce the equilibrium expression for the Haber process, assuming all components are in the gaseous phase.

2 Describe the effect of increasing temperature on the amount of each substance at equilibrium.

3 Explain the effect of temperature on the position of equilibrium. Compare this with the expected result based on Le Chatelier's principle.

4 Describe the effect of increasing the pressure on the amount of each substance at equilibrium.

5 Explain the effect of pressure on the position of equilibrium. Compare this with the expected result based on Le Chatelier's principle.

6 Describe the effect of increasing the concentration of:

  a N_2
  b H_2
  c NH_3

7 Evaluate the following claim:

Le Chatelier's principle predicts that when moles of nitrogen or hydrogen are added, reaction shifts to the right, whereas adding ammonia shifts the reaction to the left.

In your response, show that when the nitrogen/hydrogen ratio is sufficiently high, adding nitrogen shifts the reaction to the left (i.e. adding nitrogen increases the amount of hydrogen, contrary to Le Chatelier's principle). (**Hint:** Adding nitrogen when the N/H ratio is high decreases the hydrogen mole fraction.)

6.7B

PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT

How can the formation of sulfuric acid be optimised?

Practical worksheet
6.7B How can the formation of sulfuric acid be optimised?

Context

To form sulfuric acid (H_2SO_4), several reactions must take place. The first is the reaction between liquid sulfur and oxygen to form sulfur dioxide.

$$S(l)+O_2(g) \rightarrow SO_2(g) \qquad \Delta H = -297\,\text{kJ mol}^{-1}$$

The sulfur dioxide can then react with oxygen to form sulfur trioxide.

$$2SO_2(g)+O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H = -197\,\text{kJ mol}^{-1}$$

The final parts of the process involve pumping the sulfur trioxide into an absorber. Unfortunately, the direct reaction between sulfur trioxide and water to produce sulfuric acid generates large quantities of heat, vaporising the acid.

$$2SO_3(g)+H_2O(l) \rightleftharpoons H_2SO_4(g) \qquad \Delta H = -130\,\text{kJ mol}^{-1}$$

The reaction is therefore split into two. The first step is to react the sulfur trioxide with sulfuric acid to form a chemical compound called oleum ($H_2S_2O_7$).

$$SO_3(g)+H_2SO_4(l) \rightarrow H_2S_2O_7(l)$$

The oleum is then reacted with water to form two sulfuric acid molecules. This is essentially a recycling process, where one mole of acid is reacted but two are gained as products.

$$H_2S_2O_7(l)+H_2O(l) \rightarrow 2H_2SO_4(l)$$

FIGURE 1 Your company produces high quality sulfuric acid.

Aim

To develop a new process to optimise the yield of sulfuric acid

Instructions

The company you work for has identified a sudden increase in the demand for H_2SO_4. They ask you to propose a plan to modify the existing production conditions to increase the rate and yield of sulfuric acid. The company tells you that you're only able to manipulate the following reaction conditions:

- temperature
- amount of reactants/products
- pressure/volume.

Questions

1. Identify the green chemistry principles that you will need to keep in mind as you develop your proposal.
2. Identify the factors that the company is most likely to consider.
3. Explain what changes you could make to the reaction conditions to increase the rate of the reactions.
4. Identify the reaction conditions you could change to increase the yield of the reactions. Explain what you would modify and justify your response.
5. Look at the reaction conditions you have identified in questions 3 and 4 and decide which you will include in your proposal. Justify your response and refer to the trade-offs or compromises you need to make.
6. After reviewing your proposal, the company decides they want to try another approach. They give you an additional budget, which allows you to purchase any new materials or reagents you made need. Explain what you might spend this fund on and justify your response.

7.1

PRACTICAL:
CONTROLLED EXPERIMENT

How can you construct a secondary cell to electrolyse water?

Video demonstration
7.1 How can you construct a secondary cell to electrolyse water?

Practical worksheet
7.1 How can you construct a secondary cell to electrolyse water?

Risk assessment
7.1 How can you construct a secondary cell to electrolyse water?

Context

A secondary cell can operate in two modes: discharge, where reactants are converted into products to generate electrical energy, and recharge, where electrical energy is supplied to the cell to convert products back into reactants.

In this practical, you will construct a secondary cell to electrolyse water and observe the spontaneous and non-spontaneous redox reactions.

Aim

To construct a simple model of an electrolytic cell and an alkaline fuel cell

Materials

- 250 mL of 1 M KOH solution
- Power supply
- 3 wires with crocodile clips
- 100 mL measuring cylinder
- Retort stand and retort ring
- 2 semi-micro test tubes
- Voltmeter
- Plastic cup with two carbon rods embedded in the bottom and sealed with a sealant so that there are no leaks. **Note:** Ensure that the cup fits in the retort ring and does not slip through it.
- PPE

Method

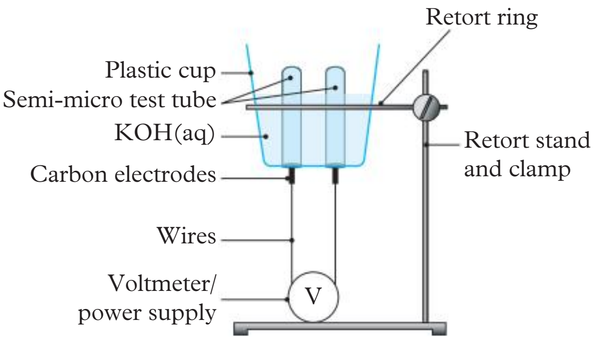


FIGURE 1 Experimental set-up of an electrolytic cell

1. Set up the retort stand with the retort ring as shown in Figure 1. Place the plastic cup fuel cell in the retort ring.
2. Pour the 1 M KOH solution into the cup so that the solution covers the electrodes.
3. Completely fill each semi-micro test tube with 1 M KOH solution and, wearing gloves, hold a finger over the top of the test tube. Turn it upside down and place it in the KOH solution and over the carbon rod. Repeat for the other electrode, ensuring that there are no bubbles in the test tubes.
4. Use the crocodile clips to connect the carbon rods to the power supply. Turn the power supply to 6 V and switch the power on for 5 minutes. **Note:** Do not allow the test tube to be completely filled with gas; there must be some KOH electrolyte left touching the carbon rods.
5. Turn the power off when the test tube containing the largest volume of gas is three-quarters full.
6. Disconnect the electrodes from the power supply and reconnect them to the voltmeter. Record the immediate voltage in Table 1 (see your obook pro).

Results

1 Draw a fully labelled diagram of the fuel cell while it is discharging, including half- and overall equations.

2 Draw a fully labelled diagram of the electrolytic cell while it is recharging, including half- and overall equations.

3 Record your results in Table 1 (see your obook pro).

Discussion

1 In electrolysis, compare the volume of gases in the test tubes. Explain your observations.

2 Explain why you shouldn't allow the gas to completely fill the electrodes.

3 Discuss the environmental issues associated with generating the gases in electrolysis. Refer to the green chemistry principles in your response.

4 Compare the theoretical E^0 value with the experimental E^0 value. Account for any difference in values.

5 Discuss the difference between the fuel cell used in this experiment and a fuel cell used in the larger scale, such as for a space shuttle.

Conclusion

Summarise the outcomes of this practical.

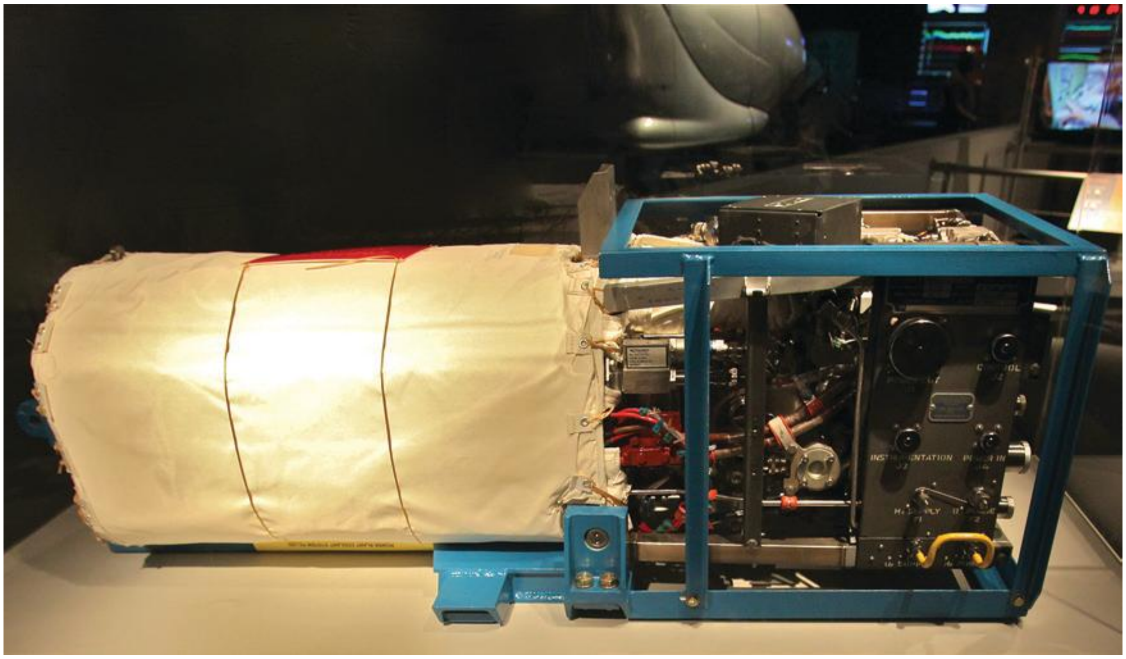


FIGURE 2 Fuel cells used in space shuttles use the same chemical principles to produce energy, but look much more complex than laboratory fuel cells.

7.2A

PRACTICAL: CONTROLLED EXPERIMENT

How can you electroplate copper onto a copper cathode?

Video demonstration
7.2A How can you electroplate copper onto a copper cathode?

Practical worksheet
7.2A How can you electroplate copper onto a copper cathode?

Risk assessment
7.2A How can you electroplate copper onto a copper cathode?

Context

Electroplating is a technique commonly used to purify metals by plating them onto an electrode.

In this practical, you will construct an electrolytic cell to plate copper onto a copper cathode.

Aim

To observe and measure the effects of electroplating copper onto a copper cathode

Materials

- 50 mL 1 M $CuSO_4$ solution
- 2 × copper strips (approx. 5 cm × 1.5 cm)
- 50 mL acetone
- Wash bottle with deionised water
- DC power supply
- 2 × 100 mL beakers
- 200 mL waste beaker
- 2 × wires with crocodile clips
- Sheet of emery paper
- Electronic balance
- Stopwatch

Method

1 Use the emery paper to clean the electrodes. Wash them with the wash bottle. Fill a 100 mL beaker with acetone and dip the electrodes into it. Allow the electrodes to air dry.

2 Weigh both electrodes by using the balance and record their mass in Table 1 (see your obook pro). Make a note of which electrodes you will use as the anode and cathode.

3 Pour 50 mL of 1 M $CuSO_4$ into a 100 mL beaker and add the electrodes.

4 Connect the cathode to the negative terminal of the power supply and the anode to the positive terminal, using the wires with crocodile clips (see Figure 1).

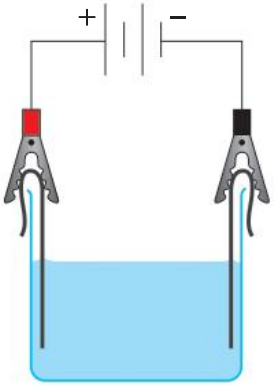


FIGURE 1 Experimental set-up of an electroplating cell

Note: It may be easier to curve the electrode around the mouth of the beaker so that it is hooked onto its side. In this way, the electrodes will not touch one another and will remain a constant distance apart.

5 Turn the settings of the power supply to 8 V. Ensure that the stopwatch is ready to start timing. Turn the power supply on and start timing for 10 minutes.

6 Turn the power supply off and disconnect the electrodes.

7 Remove the cathode and use the wash bottle to wash the copper sulfate solution remaining on the electrode into the waste beaker. Add acetone to the 100 mL beaker and dip the cathode into the beaker, allowing it to air dry.

8 Once completely dry, use the balance to weigh the electrode and record its mass in Table 1.

9 Repeat steps 7 and 8 with the anode.

Results

Record your results in Table 1 (see your obook pro).

Discussion

1 Compare the mass difference in the anode and cathode.

2 Draw a fully labelled cell diagram with half- and overall equations to demonstrate the processes occurring within the electrolytic cell.

3 Explain whether other materials could have been used as the anode or cathode. Refer to the electrochemical series in your response.

4 Explain whether another solution could have been used as the electrolyte.

Conclusion

Summarise the outcomes of this practical.

7.2B

PRACTICAL: CASE STUDY

Which is best … molten or aqueous?

Practical worksheet
7.2B Which is best … molten or aqueous?

Context

You are tasked with creating the best method for the electrolysis of various mixtures.

In this practical, you will design electrolytic cells to achieve the aims indicated in Table 1.

TABLE 1 Chemical mixtures available and purposes of the electrolytic cell

Mixtures available	Aims
NaCl, KF, $MnCl_2$	To generate manganese metal at the cathode without releasing toxic chlorine into the laboratory
Cu, $Al(NO_3)_3$, $Zn(NO_3)_2$	To generate zinc metal at the cathode of a cell
MnF_2, H_2S	To generate solid sulfur in a manner that will prevent the sulfur from spontaneously reacting with water

FIGURE 1 Molten liquids are melted down solids, while aqueous solutions are solids that have been dissolved in water.

Aim

To design the most appropriate electrolytic cells for the described purposes

Questions

For each mixture:

1 Draw the molten electrolytic cell for all three mixtures, ensuring that you include and label: an appropriate anode and cathode material, electrode polarity, oxidation and reduction half-equations, the strongest oxidising and reducing agents, the overall equation, the movement of ions in the cell, the movement of electrons and the voltage which must be applied to the cell.

2 Draw the aqueous electrolytic cell for all three mixtures, ensuring that you include and label: the appropriate anode and cathode material, electrode polarity, oxidation and reduction half-equations, the strongest oxidising and reducing agents, the overall equation, the movement of ions in the cell, the movement of electrons and the voltage which must be applied to the cell.

3 Explain what would happen to the aqueous cell if it was run at non-standard conditions. Include appropriate half-equations in your answer.

4 Describe the recommendations you would make to the student regarding the best cell to use and the conditions that it should be run with.

For each cell:

5 Compare the ‘best’ cell for each mixture. Identify whether they were all the same type (molten or aqueous). Explain why or why not.

6 Explain one advantage and one disadvantage to using molten and aqueous electrolytic cells.

7 Explain why having a mixture of molten salts may be both an advantage and a disadvantage in a molten electrolytic cell.

8 Explain two major purposes of a porous membrane in an electrolytic cell.

9 Explain the structural requirements to ensure that a secondary cell is rechargeable.

8.1 PRACTICAL: CLASSIFICATION & IDENTIFICATION

How can the physical properties of organic compounds be used to classify them?

Practical worksheet
8.1 How can the physical properties of organic compounds be used to classify them?

Context

What determines the physical properties of organic compounds? You have learnt that it is the presence of different functional groups that affects the boiling and melting points and solubility of organic compounds.

In this practical, you will apply your knowledge to identify organic compounds based on their physical and chemical properties.

FIGURE 1 The boiling point is an important physical property.

Aim

To identify organic chemicals based on their physical and chemical properties

Questions

1 Draw the structural formulas of the following molecules onto small cards so that you can shuffle them about and stick them in the table when you have correctly identified them:
  - methanoic acid
  - ethanol
  - ethanoic acid
  - propane
  - propene
  - 1-chloropropane
  - 2-chloropropane
  - propan-1-ol
  - propanal
  - propanone
  - propan-1-amine
  - propanoic acid
  - propanamide
  - ethyl methanoate
  - butanal
  - butanone
  - pentane
  - butan-1-ol
  - butan-2-ol

2 Copy out Table 1 into your logbook, leaving space in the first column for the molecules you have drawn.

TABLE 1 Physical properties of compounds A–S

Compound	Melting point (°C)	Boiling point (°C)	Solubility in water	Strongest intermolecular bonding between molecules
A	–187.6	–42.1	Insoluble	Dispersion forces
B	–129.7	36.1	Insoluble	Dispersion forces
C	–185.0	–47.7	Insoluble	Dispersion forces
D	–80.0	53.0	Slightly soluble	Dipole–dipole
E	–81.0	48.0	Soluble	Dipole–dipole
F	–123.0	46.5	Insoluble	Dipole–dipole
G	–96.0	75.0	Soluble	Dipole–dipole
H	–118.0	35.0	Insoluble	Dipole–dipole
I	–87.0	80.0	Soluble	Dipole–dipole
J	–94.0	56.0	Soluble	Dipole–dipole
K	–127.0	97.0	Soluble	Hydrogen bonding
L	–23.5	141.0	Soluble	Hydrogen bonding
M	8.3	100.5	Soluble	Hydrogen bonding
N	–90.0	117.0	Soluble	Hydrogen bonding
O	–83.0	48.0	Soluble	Hydrogen bonding
P	16.2	117.5	Soluble	Hydrogen bonding
Q	–114.0	78.0	Soluble	Hydrogen bonding
R	77.5	213.0	Soluble	Hydrogen bonding
S	–115.0	98.0	Soluble	Hydrogen bonding

3 Using your knowledge of the physical properties of organic compounds, match the molecules in the list with those in Table 1.

4 Consider the following two pairs of compounds: N and S; F and H. The pairs have the same molar mass, but different physical properties. Explain the difference in their physical properties.

5 Predict the melting and boiling points for the following compounds. Sketching a quick graph may help.

- **a** Butane
- **b** Hexane
- **c** Butanoic acid
- **d** Ethanal
- **e** Pentanal
- **f** Pentan-1-ol

6 Explain why carboxylic acids have high boiling points.

7 Identify any molecules that were particularly difficult to match to their properties. Discuss what made it difficult.

8.4
PRACTICAL: MODELLING

How do the shapes of molecules affect their intermolecular bonding and physical properties?

Video demonstration
8.4 How do the shapes of molecules affect their intermolecular bonding and physical properties?

Practical worksheet
8.4 How do the shapes of molecules affect their intermolecular bonding and physical properties?

Risk assessment
8.4 How do the shapes of molecules affect their intermolecular bonding and physical properties?

Context

As you have learnt in Chapter 8, carbon–carbon double bonds and branches affect the shapes of molecules. This will affect the physical properties of molecules.

In this practical, you will use a molecular modelling kit to create the molecules heptane and hep-1-tene, along with their isomers, and discuss how the branches and double bonds will affect the physical properties.

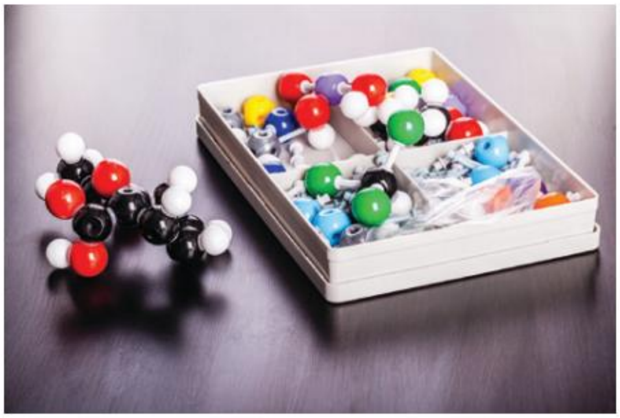


FIGURE 1 Molecular modelling kits can be used to determine the shapes of molecules.

Aim

To model the shapes and structures of various organic chemicals, identify the intermolecular forces and determine how they affect the physical properties

Materials

- Molecular modelling kit

Method

Part A – Effect of branching

1 Use the molecular modelling kit to make the following molecules:
  - heptane (C_7H_{16})
  - isomers of heptane:
    2-methylhexane
    3-methylhexane
    2,3-dimethylpentane
    2,4-dimethylpentane
    3-ethylpentane
    2,2,3-trimethylbutane

Part B – Effect of the double bond position

2 Use the molecular modelling kit to make the following molecules:
  - positional isomers of heptene (C_7H_{14}):
    hept-1-ene
    hept-2-ene
    hept-3-ene

3 Draw each of the structures into your logbook, clearly showing the shape of each of the molecules.

Questions

1 Discuss the impact of the shape of the molecules on the strength of the intermolecular forces that occur between the same molecules, using examples from the molecules that you have created.

2 Discuss how the shape of each of the isomers of heptane and heptene will affect the following properties:
  a boiling point
  b melting point
  c viscosity.

3 Predict the order of the boiling points of the molecules that you have created in Part A.

4 Predict the order of the boiling points of the molecules that you have created in Part B.

5 Research the boiling points of each of the molecules you have created. Explain how they correspond to the trends that you predicted in questions 3–4.

9.1A

PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT

How can you produce an ester?

Practical worksheet
9.1A How can you produce an ester?

Context

A chemist wants to produce the ester, ethyl butanoate.

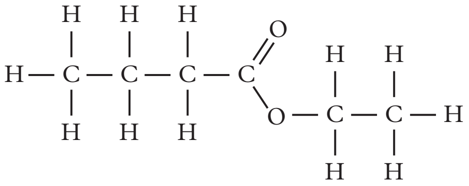


FIGURE 1 Ethyl butanoate

They know that a series of reactions is required to produce the ester but have no idea where to start. The chemist approaches you for your help.

The chemist has access to the following:

- reactants: H_2O, NaOH, KOH, H_2, HCl, Cl_2, HBr, Br_2, NH_3
- alkanes: methane, ethane, propane, butane
- alkenes: ethene, propene, but-1-ene, but-2-ene
- catalysts: $H^+/Cr_2O_7^{2-}$, H_3PO_4, $H^+/KMnO_4$, Cl_2, H_2SO_4, Ni, palladium-on-charcoal, enzymes, UV light
- an oven.

In this practical, you will develop a process for the synthesis of ethyl butanoate.

Aim

To develop a process for the production of ethyl butanoate

Instructions

1 Draw the structure of ethyl butanoate.

2 Identify the two reactants that you will need to react together to form ethyl butanoate. Provide their full chemical structures and IUPAC names.

3 Select your starting materials from the alkanes and alkenes available. Justify your choices.

4 For each starting material, design a pathway to produce the chemicals you identified in Question 2. For each step, include the:
   - name of the reaction
   - reactants required
   - catalysts required
   - temperature required (if different to room temperature)
   - molecular formula, structural formula and IUPAC name of the compound formed.

5 You now have the two chemicals you need to make ethyl butanoate. Describe how you would use these two reactants to produce the final ester. Name the reaction and the catalysts required.

6 Combine all of the above to create a flowchart to show the chemist.

Questions

1 Discuss the yield of ethyl butanoate product you would expect to achieve from this process. Explain why the yield is unlikely to be 100%.

2 Discuss any challenges you experienced while designing the process. Were there any processes for which you had trouble deciding on a reactant or catalyst? Explain how you made your final choices.

3 **a** Analyse the process you developed and identify any by-products.

  b Discuss why it is important to be aware of by-products. Refer to the green chemistry principles in your response.

4 Explain whether it is possible to modify the process to produce 'greener' ethyl butanoate. Suggest any changes you could make to develop a more sustainable product.

5 Conduct a risk assessment for the reactions in your process. Highlight any key hazards and provide some instructions to the chemist for how to minimise risks.

6 Discuss the sociocultural, economic, political, legal and/or ethical factors that industry professionals would need to consider when designing processes to produce chemicals like ethyl butanoate.

9.1B
PRACTICAL: CONTROLLED EXPERIMENT

How do you create artificial scents?

Video demonstration
9.1B How do you create artificial scents?

Practical worksheet
9.1B How do you create artificial scents?

Risk assessment
9.1B How do you create artificial scents?

Context

Chemists can create artificial scents for foods, perfumes and medicines. These are known as esters. An ester is created by adding an alcohol and carboxylic acid together in the presence of sulfuric acid.

In this practical, you will investigate the different scents created by esters and compare them with theoretical scents.

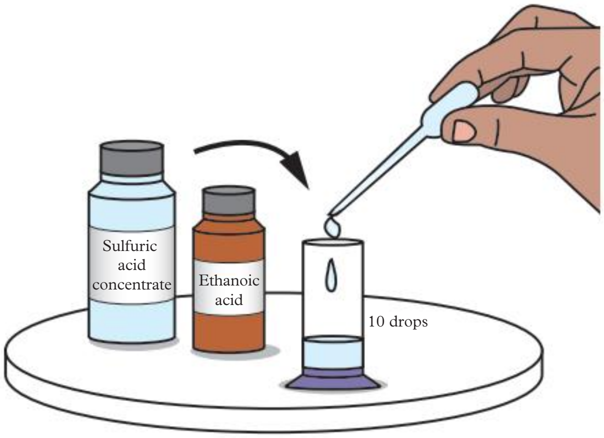


FIGURE 1 Preparation of an ester involves mixing an alcohol and a carboxylic acid.

Aim

To prepare several esters, compare their theoretical scents and actual scents, and understand the naming of esters and the reaction pathways that form them

Materials

- Concentrated sulfuric acid (H_2SO_4)
- Pasteur pipettes
- Methanol
- Propan-1-ol
- Ethanoic acid
- Octanoic acid
- Boiling water
- 60 mL of cold deionised water
- 6 × micro test tubes
- 1 × 250 mL beaker
- 6 × 50 mL beakers
- A kettle or Bunsen burner and retort stand set-up

Method

1 Label the six micro test tubes A–F.

2 Add 10 drops of the selected alcohol and 10 drops of the carboxylic acid into test tubes, as shown in Table 1.

TABLE 1 Alcohol and carboxylic acid mixtures

A	B	C	D	E	F
Methanol and ethanoic acid	Ethanol and ethanoic acid	Propan-1-ol and ethanoic acid	Methanol and octanoic acid	Ethanol and octanoic acid	Propan-1-ol and octanoic acid

3 Add two drops of concentrated sulfuric acid into each of the test tubes.
4 Put the test tubes into a beaker containing the boiling water. Make sure the top of each micro test tube is above the water line and leave to brew for 10 minutes.
5 Set out the six 50 mL beakers, each containing approximately 10 mL cold deionised water.
6 Gently shake the test tube and then pour it over the cold water.
7 Waft the scent created and record your observations.

Results

Record your results in Table 2 (see your obook pro). Compare the scents of the esters you created with the expected scent created by the alcohol and carboxylic acid pair.

Discussion

1 Explain the role of concentrated sulfuric acid in these reactions.
2 In your logbook, draw the full reaction for each of the esters you have created. Use full structural formulas and correct IUPAC naming for each.
3 Identify any esters where the scent you created differed from the theoretical scent. Suggest a reason for these differences.
4 When the ester is poured over the cold water, you will notice it sits on top of the water. Explain why esters are generally not soluble in water.

Conclusion

Fill in the blanks to complete the conclusion:

> Artificial scents are created by adding an ____________ to a ______________ and reacting them in the presence of a ______________ and _____________. They react to form an _______________ functional group and a _________________.

9.3

PRACTICAL:
CONTROLLED EXPERIMENT

What is the percentage yield of aspirin?

Video demonstration
9.3 What is the percentage yield of aspirin?

Practical worksheet
9.3 What is the percentage yield of aspirin?

Risk assessment
9.3 What is the percentage yield of aspirin?

Context

Chemists use a range of reaction pathways to create a variety of chemicals, including medicines.

In this practical, you will synthesise aspirin from salicylic acid and acetic anhydride, and calculate the yield of aspirin from the reaction.

Aim

To synthesise aspirin and calculate the percentage yield

Materials

- 3 g salicylic acid
- 10 mL acetic anhydride
- 1 mL concentrated sulfuric acid
- 250 mL conical flask
- 10 mL measuring cylinder
- Weigh boat
- Spatula
- Filter flask
- Large watch glass
- Glass rod
- Buchner funnel
- Filter paper
- Ice water bath
- Analytical balance
- PPE

Method

1 Weigh approximately 3 g of salicylic acid in a weigh boat and record the mass.

2 Transfer the mass into a clean conical flask.

3 In the fume hood, add 10 mL of acetic anhydride and 1 mL of sulfuric acid to the conical flask. Swirl the mixture until everything has dissolved. Do not use heat to dissolve.

4 Add 100 mL of ice water to the conical flask.

5 The acetylsalicylic acid should begin to crystallise out, but to assist the process, scratch the walls of the flask with a glass rod. You can keep cooling the mixture slightly in an ice bath until crystallisation is complete. The product will appear as a solid mass when crystallisation is complete.

6 Record the mass of your filter paper and watch glass.

7 Vacuum filter the product using a Buchner funnel, as shown in Figure 1.

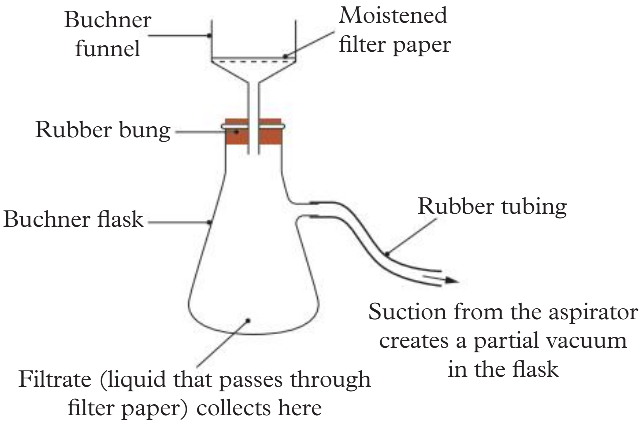


FIGURE 1 Buchner funnel set-up

8 Rinse the conical flask with deionised water to make sure all of the product is collected.

9 Once all of the product has been collected, allow the crystals and filter paper to dry and re-weigh them.

Results

1 Record the:

   a initial mass of the salicylic acid

   b initial mass of the filter paper and watch glass c final mass of the dried aspirin + filter paper + watch glass.

2 Calculate the final mass of the product.

  Final mass of product = (final mass of the dried aspirin + filter paper + watch glass) – (initial mass of the filter paper and watch glass)

Discussion

1 Explain the role of the concentrated sulfuric acid in this experiment.

2 Calculate which reactant is limiting if the density of acetic anhydride is 1.08 $g\,mL^{-1}$.

3 Calculate the theoretical mass of aspirin that could be obtained from this experiment, using your original mass of salicylic acid.

4 Use your answer to Question 3 to calculate the percentage yield of aspirin.

5 Suggest alternative reactants that could be used to produce the same aspirin molecule, but result in an increased atom economy.

Inquiry: How do you increase the percentage yield?

1 Select an independent variable from the list below that you will change to increase the percentage yield of aspirin:
  - volume of acetic anhydride
  - mass of salicylic acid
  - temperature of the solution when dissolving in Step 3
  - change in scratching technique
  - seeding aspirin crystals instead of scratching
  - change in filtration technique.

2 Write a hypothesis for your investigation.

3 Write a method to test your hypothesis.

4 Repeat the experiment and calculate the new percentage yield.

5 Compare your original percentage yield and your new percentage yield.

Conclusion

Write a suitable conclusion for this experiment.

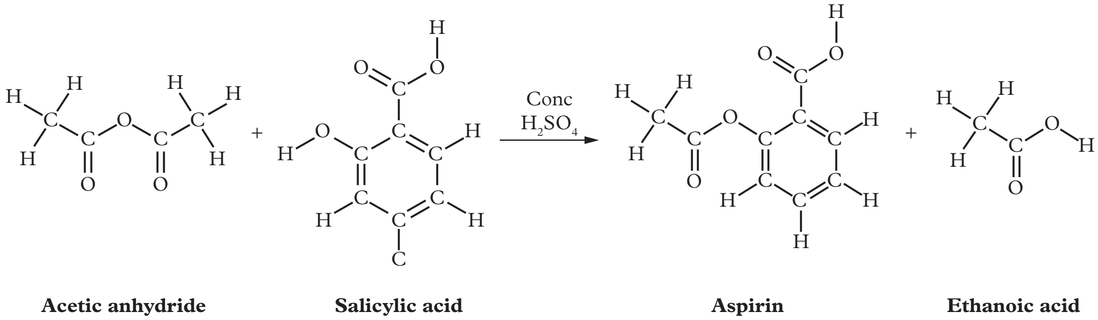


FIGURE 2 Pathway for synthesis of aspirin

10.1
PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT

How are qualitative tests used to identify functional groups?

Practical worksheet
10.1 How are qualitative tests used to identify functional groups?

Context

The functionality of organic molecules can be determined using simple qualitative tests. These tests can determine whether a molecule is unsaturated, containing one or more carbon–carbon double or triple bonds, or has a functional group.

In this practical, you will explore a range of qualitative tests to identify functional groups. The tests and their underlying principles are described below:

- Testing for alcohols: Primary and secondary alcohols will undergo oxidation when reacted with an oxidising agent such as acidified potassium permanganate.
- Testing for carboxylic acids: Metal carbonates will produce carbon dioxide gas when reacted with acids.
- Testing for unsaturation: Bromine can be added to an unsaturated hydrocarbon to break the carbon–carbon double bond in an alkene or a carbon–carbon triple bond in an alkyne.
- Testing for polarity: The ‘like with like’ rule is used to determine the solubility of substances. Substances with similar properties will be miscible. For example, a non-polar substance will dissolve in non-polar solvents. Likewise, any molecule or compound with whole or partial charges, such as ionic salts and polar molecules, will dissolve in substances which also have whole or partial charges, such as water. If they are immiscible, they will form layers, with the densest layer at the bottom.

FIGURE 1 A non-polar substance is not miscible with a polar substance.

Aim

To develop and test a procedure to determine the functional groups in organic molecules using qualitative tests

Instructions

Develop processes to determine whether a molecule is:

a an alcohol
b a carboxylic acid
c unsaturated
d polar.

Each test should have its own individual process. These processes must include instructions for how to perform the test, the chemicals required, an answer key to explain what should be observed if the test is positive or negative, and what inferences can be made from the results.

If possible, your teacher will give you an unknown organic substance that you can use to perform your tests on.

Questions

1. For the alcohol tests, explain whether all types of alcohols will experience a positive test.
2. For the carboxylic acid tests, explain why bubbles will be generated when reacted with a base. Use a chemical equation to support your answer.
3. Identify any other tests that could be used to identify alcohols and carboxylic acids. Explain the results that you would expect in these tests.
4. For the unsaturation tests, explain the results that you would expect to observe and how they are able to justify the presence or absence of a carbon–carbon double bond. Include a chemical equation to support your answer.
5. For the polarity tests, explain the results that you would expect to observe and how they are able to justify the polarity of organic compounds. Your response must include an explanation of the intermolecular interactions between molecules.
6. Explain why some organic molecules will form layers when placed in water and what this tells you about the identity of the molecule. Your response must include an explanation of the intermolecular interactions between molecules.
7. A student is asked to characterise an organic molecule based on whether it:
   - is saturated or unsaturated
   - contains a hydroxyl group (is an alcohol)
   - contains a carboxyl group (is a carboxylic acid)
   - is polar or non-polar.

   The unknown molecule that they are given is a sample of ethanoic acid. They follow the procedures you have designed in this practical. Predict their results and indicate what inferences can be made based on their results.

FIGURE 2 When carboxylic acids react with metal carbonates, carbon dioxide gas bubbles are produced.

10.4

PRACTICAL: CONTROLLED EXPERIMENT

What is the concentration of vitamin C in orange juice?

Video demonstration
10.4 What is the concentration of vitamin C in orange juice?

Practical worksheet
10.4 What is the concentration of vitamin C in orange juice?

Risk assessment
10.4 What is the concentration of vitamin C in orange juice?

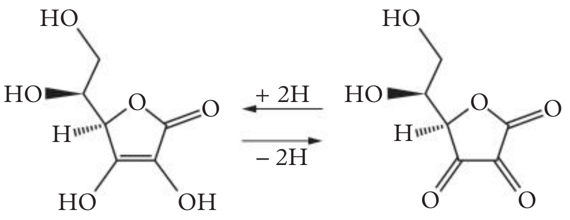
Context

Quality control chemists use redox titrations to confirm the concentrations of specific molecules within a food product. Ascorbic acid (vitamin C) content in orange juice can be quantitatively determined using this method. The ascorbic acid is oxidised to form dehydroascorbic acid in the presence of an oxidising agent. In this reaction, the oxidising agent used is a solution of iodine, which undergoes reduction to form a solution of iodide. Iodine therefore acts as the oxidising agent in this reaction.

$$C_6H_8O_6(aq) \rightarrow C_6H_6O_6(aq) + 2H^+(aq) + 2e^-$$
$$I_2(aq) + 2e^- \rightarrow 2I^-(aq)$$

The end point of the reaction is reached when all of the ascorbic acid has reacted and there is an excess of iodine in the flask. This end point is visualised using a starch indicator, which turns a dark blue–black in the presence of iodine.

In this practical, you will experimentally determine the concentration of vitamin C in orange juice.



Ascorbic acid **Dehydroascorbic acid**

FIGURE 1 The oxidation of ascorbic acid

Aim

To determine the concentration of vitamin C in orange juice and compare it to the concentration stated on the food label

Materials

- Retort stand and clamp
- White tile
- Three 250 mL conical flasks
- 50 mL burette
- Small glass funnel
- 250 mL waste beaker
- Three 100 mL transfer beakers
- 20 mL pipette. **Note:** This may need to change based on the concentration of the ascorbic acid.
- Pipette bulb
- Wash bottle
- Pulp-free orange juice – enough for 3 concordant titres. **Note:** The orange juice can be replaced by any product containing ascorbic acid, such as vitamin C tablets or lemon juice. In the case of the tablets, their mass must be determined first before they are crushed in the conical flask and dissolved with 20 mL of water.
- 0.005 M iodine solution
- Starch indicator in a dropper bottle

Method

Part A – Pre-lab

1 Look at the label on the orange juice packaging and record its vitamin C concentration.

2 Calculate the titre volume of iodine, based on the theoretical concentration of ascorbic acid from the product label. Determine whether a 20 mL aliquot of juice is adequate.

Part B – Experiment

3 Set up the retort stand and clamp with the white tile on the base of the retort stand.

4 Add some water from the wash bottle to one of the 100 mL transfer beakers. Use this water to rinse the burette and pipette into the waste beaker.

5 Add some of the iodine solution to the second 100 mL transfer beaker. Put the funnel in the top of the burette and add some iodine solution to wash the burette. Open the tap to allow the now diluted iodine solution to drain into the waste beaker.
6 Using the funnel again, add iodine solution to the burette with the tap open. Once you are sure that no bubbles have formed below the tap of the burette, close the tap and continue to fill the burette as close to the 0.00 mL line as you can, but not over it. Remove the funnel and then record the initial volume of the burette in your logbook.
7 Add some of the orange juice to the third 100 mL transfer pipette and use it to rinse the pipette. Dispose of this into the waste beaker.
8 Using the pipette, measure a 20.00 mL aliquot of orange juice and dispense it into the conical flask.
9 Add approximately five drops of the starch indicator to the orange juice.
10 Titrate the orange juice with the iodine solution until the first permanent pale blue colour is observed. If you are unsure of the end point, record the final titre volume of the burette in your logbook and then add one more drop of iodine. If it turns a darker blue, you have already reached the end point.
11 Repeat the experiment using new 20.00 mL aliquots of orange juice in the remaining conical flasks until you have recorded three concordant titres. Fill the burette between each titration, if necessary; you should not be refilling a burette in the middle of a titration. Remember that concordant titres cannot differ by more than 0.1 mL.

Results

1 Construct a results table which includes the initial and final burette volume and the titre volume of each titration.
2 Calculate the average titre volume of iodine based on the concordant titres.
3 Calculate the concentration of vitamin C in orange juice by following the steps:
   a Calculate the amount, in mol, of iodine.
   b Write a balanced chemical equation for the reaction, including states.
   c Calculate the amount, in mol, of ascorbic acid.
   d Calculate the concentration of ascorbic acid in the same units that appear on the orange juice bottle.

Discussion

1 Compare the experimental concentration that you have calculated to the theoretical concentration on the orange juice bottle.
2 In the experiment, a student does not wash the burette with iodine before filling it to start the titration. Predict the effect that this will have on the calculated concentration of vitamin C.
3 When reading the burette, a student measures the volume from below eye level. Predict the effect that this will have on the titre volume and, therefore, the final concentration of vitamin C.
4 Justify whether the data is accurate, precise, repeatable and reproducible.
5 Evaluate whether this experiment is valid. Justify your response.

Inquiry: What if the concentration of iodine was halved?

1 Write a hypothesis for your investigation.
2 Identify the independent variable that you will need to change to test your hypothesis.
3 Describe the changes you will need to make to the original experimental method.
4 Predict the titre volume you would expect from your experiment. Justify your response.
5 Calculate the expected concentration of vitamin C.
6 Compare the size of the titres from your initial experiment to the titres from your inquiry. Use your understanding of systematic errors and resolution of glassware to comment on the accuracy and precision of the data.

Conclusion

Write a suitable conclusion based on the experimental and theoretical concentrations of vitamin C in orange juice.

11.5A

PRACTICAL:
CASE STUDY

How can analytical techniques be used to differentiate between isomers?

Practical worksheet
11.5A How can analytical techniques be used to differentiate between isomers?

Context

Your job is to outline how a chemist would use IR, ^{1}H-NMR and ^{13}C-NMR to identify a chemical in a sample. You know that the organic molecule has the formula $C_3H_6O_2$ and that it can exist as a carboxylic acid or an ester.

In this practical, you will predict the results of spectroscopy analysis and provide advice to the chemist.

FIGURE 1 ^{1}H-NMR spectra can be used to characterise organic molecules.

Aim

To outline how the principles of ^{1}H-NMR, ^{13}C-NMR and IR can lead to the different spectra for a carboxylic acid and ester isomer, and how you could use them to differentiate between the chemicals

Questions

1 There are a few isomers of $C_3H_6O_2$.
   - **a** Draw the three isomers of $C_3H_6O_2$. There is one carboxylic acid isomer and two ester isomers.
   - **b** Identify the IUPAC names of each isomer.

2 Describe what the IR spectra would look like for both an ester and a carboxylic acid. Outline the features that each spectrum would have that would help you determine if $C_3H_6O_2$ was an ester or a carboxylic acid.

3 ^{1}H-NMR is used to determine the number of hydrogen environments.
   - **a** Draw a high-resolution ^{1}H-NMR spectrum for the carboxylic acid isomer of $C_3H_6O_2$. Include the chemical shift and splitting patterns.
   - **b** There are two isomers of $C_3H_6O_2$ that could exist as esters. Draw the high-resolution ^{1}H-NMR spectra for both, including the chemical shift and splitting patterns.
   - **c** Discuss how the ^{1}H-NMR spectra differ for the two ester isomers. Discuss whether you are able to use ^{1}H-NMR to differentiate between the ester isomers of $C_3H_6O_2$.
   - **d** Discuss whether there is enough difference in the three ^{1}H-NMR spectra for you to identify the chemical.

4 Draw the ^{13}C-NMR spectra for the carboxylic acid and the ester isomers. Discuss whether there is enough difference in the three ^{13}C-NMR spectra for you to identify the chemical.

5 Discuss whether you could use just one of the analytical techniques to identify the isomer of $C_3H_6O_2$. Explain whether you would need to combine more than one technique to identify the chemical.

6 Discuss whether mass spectroscopy would be of any use in this situation.

7 Develop a flowchart to help the chemist identify the isomer using the spectra generated.

11.5B

PRACTICAL:
CASE STUDY

Lab clean-up duty – what's in the bottle?

Practical worksheet
11.5B Lab clean-up duty – what's in the bottle?

Context

You have been employed to work in a research lab and it is your first day. Unfortunately, when you arrive at work, you find that the lab is a mess! There are unlabelled bottles of chemicals everywhere.

FIGURE 1 Your new lab is a mess.

Your first task is to clean up. Instead of throwing everything out and generating volumes and volumes of chemical waste, you decide to put to use the analytical chemistry skills you learnt in VCE Chemistry. Luckily for you, there is a mass spectrometer, infrared spectrometer and hydrogen and carbon NMR spectrometer in the lab that you can use.

You start with a shelf labelled 'Ketones, aldehydes and alcohols', in the hopes that there is some semblance of organisation already. You grab four bottles off the shelf and take them over to be tested. In this practical, you will analyse the spectroscopy data from your tests and identify the compounds in the bottles.

Aim

To identify the contents of four unlabelled chemical bottles using a variety of analytical techniques

Questions

The IR, MS, ^{1}H NMR and ^{13}C NMR spectra for samples obtained from the four bottles are shown below. They are labelled A–D, shown over the next three pages.

Discussion

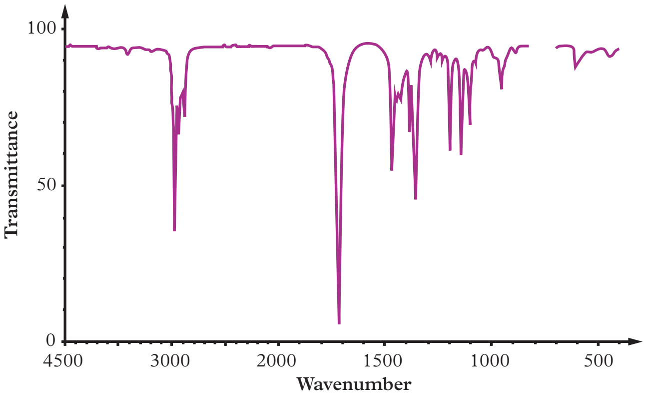
1 Identify the molecular ion peak in each of the mass spectra. Explain what this tells you about each chemical A–D.
2 Draw the full chemical structures for chemicals A–D.
3 Identify the IUPAC names of each of the chemicals A–D.
4 Discuss whether running both ^{13}C-NMR and ^{1}H-NMR was a good idea for all four chemicals.
5 Explain whether it was useful to use IR for all four chemicals. Discuss the information you got from each IR.
6 Identify which of the four analytical techniques gave you the most useful information. Justify your response.

Conclusion

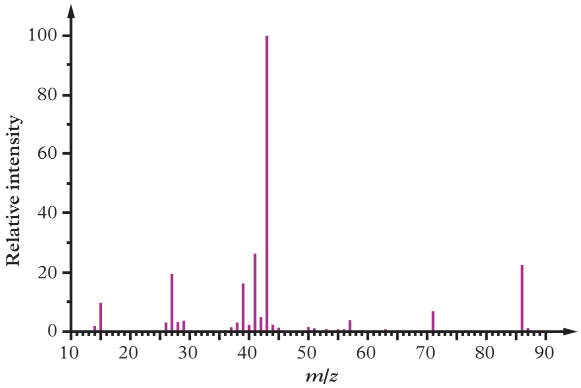
Identify the compounds in each of the four unknown bottles and write a conclusion addressing whether you were able to achieve your aim for this practical.

Sample A

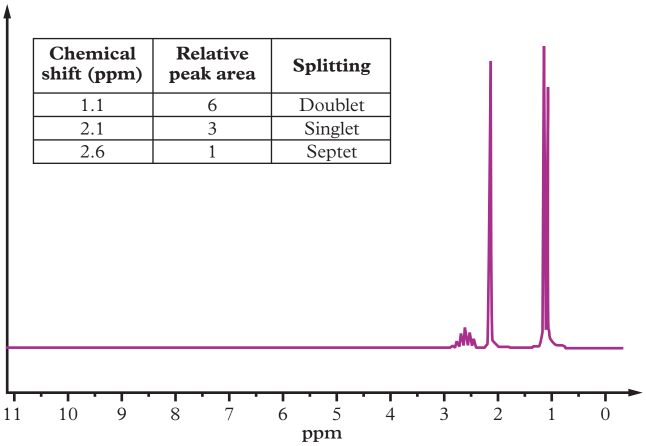
IR spectrum



Mass spectrum

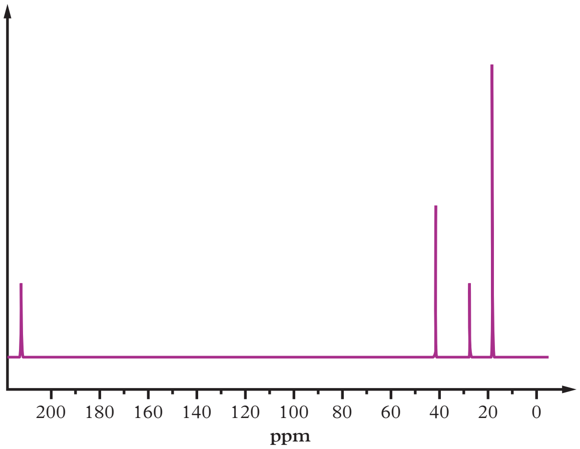


^{1}H-NMR spectrum



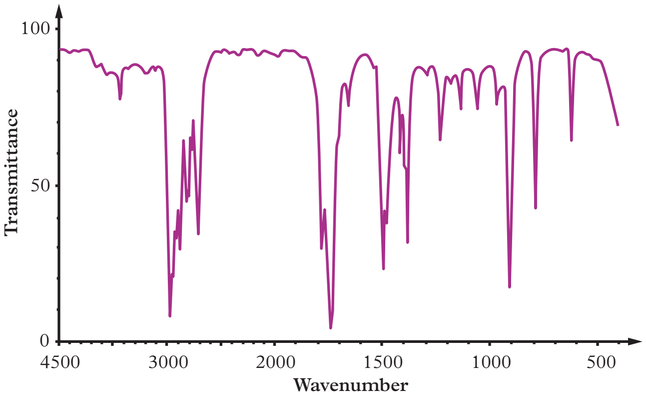
Chemical shift (ppm)	Relative peak area	Splitting
1.1	6	Doublet
2.1	3	Singlet
2.6	1	Septet

^{13}C-NMR spectrum

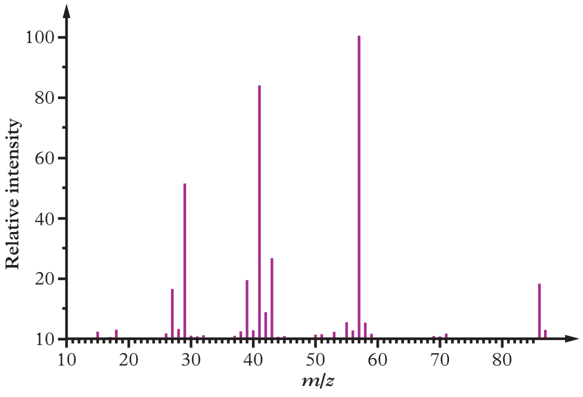


Sample B

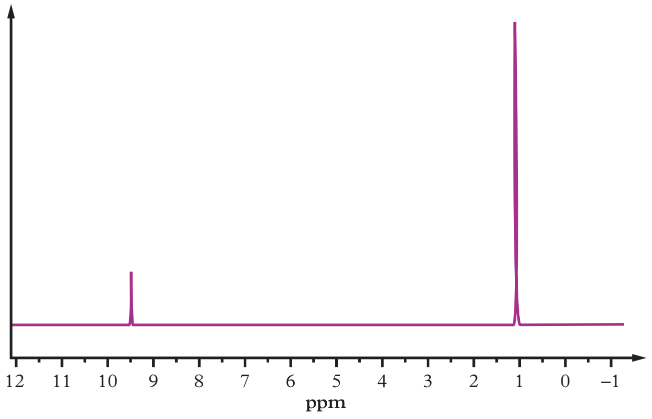
IR spectrum



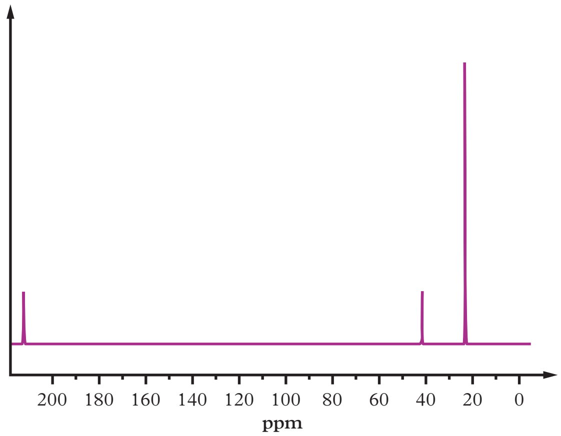
Mass spectrum



^{1}H-NMR spectrum

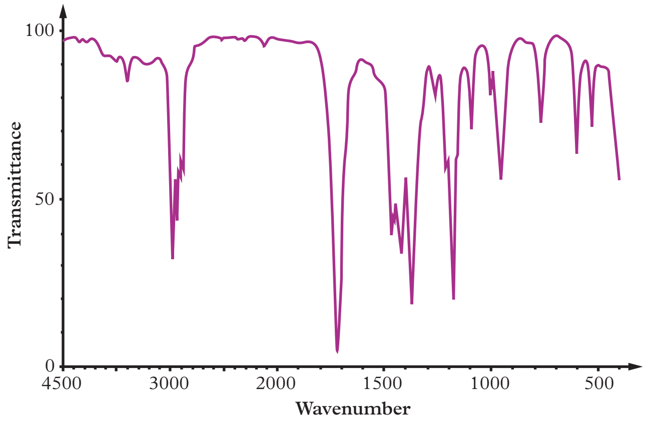


^{13}C-NMR spectrum

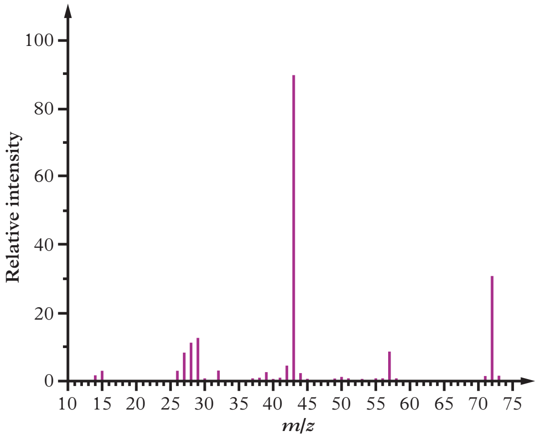


Sample C

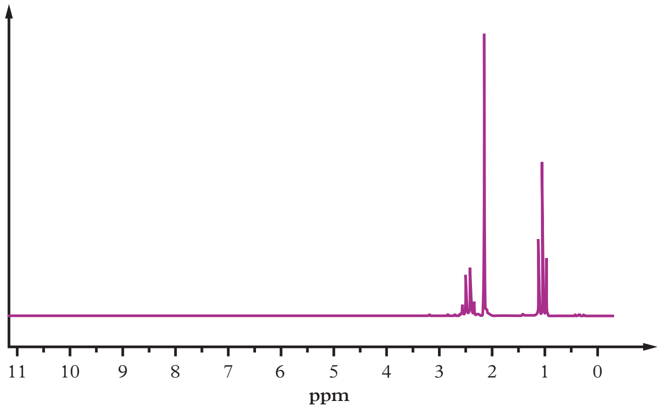
IR spectrum



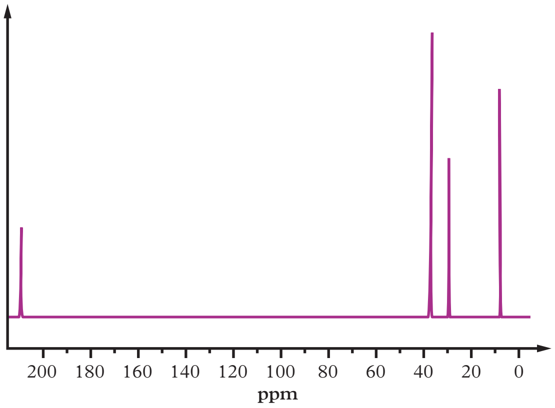
Mass spectrum



^{1}H-NMR spectrum

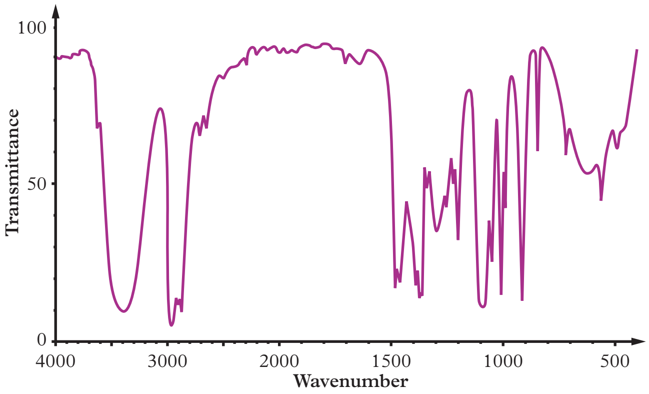


^{13}C-NMR spectrum

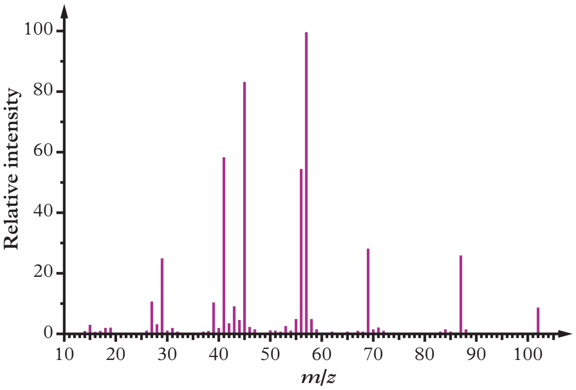


Sample D

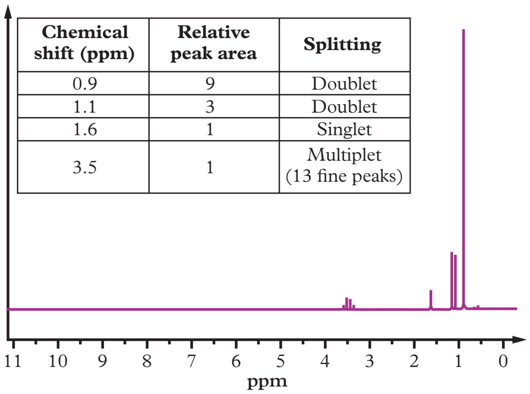
IR spectrum



Mass spectrum

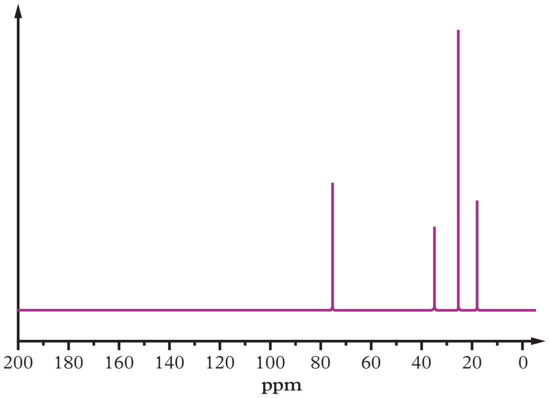


^{1}H-NMR spectrum



Chemical shift (ppm)	Relative peak area	Splitting
0.9	9	Doublet
1.1	3	Doublet
1.6	1	Singlet
3.5	1	Multiplet (13 fine peaks)

^{13}C-NMR spectrum



11.7

PRACTICAL: FIELDWORK

How is instrumental analysis used in industry?

Practical worksheet
11.7 How is instrumental analysis used in industry?

Context

Instrumental analysis is a core part of product development and quality control in industry. Analytical chemistry is used to qualitatively and quantitatively analyse organic compounds to make sure the correct product is made, that it is free of contaminants, and that the correct amount of the chemical is present. In this practical, you will investigate analytical techniques used in industry.

FIGURE 1 Samples loaded into tubes for NMR analysis

Aim

To visit a company that produces chemical products to gain a better understanding about real-life applications of analytical techniques

Instructions

Visit a company that produces chemical products. Some examples are indicated in Table 1.

TABLE 1 Example facilities to visit

Industry	Potential facilities
Paint	Dulux
Cosmetics	Aesop, L'Oréal, Ego, Intraceuticals
Food products	Yakult, Carlton United Breweries, Nestlé
Pharmaceuticals	Baxter, Pfizer, Acrux, CSL, Aspen Pharma
Agrochemicals	Eureka

Questions

1 Describe the type of product produced by the company.

2 Identify which types of instrumental analysis you learnt about in Chapter 11 are also used by the company.

3 Because you are visiting a company that develops products for commercial use, the chemicals used in their products may be protected by intellectual property and information you will gain about these chemicals may be limited. Instead, you may like to explore analysis of contaminants.

   a Identify a contaminant that is routinely tested for using IR, ^{1}H-NMR, ^{13}C-NMR or MS.

   b Explain why it is considered a contaminant and identify the effects it may have on the product if not removed or reduced.

   c Describe the results you would expect to obtain if the contaminant was present in the sample of product; i.e. what would the spectrum look like?

   d Explain if the analysis of contaminants is qualitative or quantitative. Justify your response.

   e If possible, obtain a sample spectrum from the scientists. Analyse the spectrum, to evaluate whether contaminants are present.

4 Explain how scientists working in industry make sure that their testing is valid. Compare this with what you have been taught in VCE Chemistry.

5 Describe some other types of analysis that are used by the company and explain why these tests are necessary.

6 Outline the implications of not conducting instrumental analysis of organic products.

12.1 PRACTICAL: FIELDWORK

How can you extract active ingredients from medicinal plants?

Practical worksheet
12.1 How can you extract active ingredients from medicinal plants?

Context

Aboriginal and Torres Strait Islander Peoples have been using medicinal plants for thousands of years. The use of plants for therapeutic purposes is deeply embedded in Indigenous culture and traditional knowledge, which has been passed down through generations.

Plants such as the tea tree, eucalyptus and kakadu plums are used for their medicinal properties. In some cases, the plants are prepared like tea, while in others they are used as a topical treatment. These medicines are still used today, and many, particularly tea tree and Eucalyptus, have become well-known in the wider community.

In this practical, you will collect medicinal plants from the environment and develop a process to extract medicinal compounds from them.

FIGURE 1 Hop bush is an Australian plant that has medicinal properties.

Aim

To collect medicinal plants from their habitat and extract their active ingredients by distillation

Instructions

Visit a nearby location that has native medicinal plants. Species of eucalyptus tree are quite easy to find in many parts of Victoria. Harvest a handful of leaves (without twigs or branches) and obtain a few millilitres of plant oil by steam distillation.

Questions

1 State the traditional name and Latin name for the species you obtained.
2 List the ailments that this plant is used to treat.
3 List the active ingredients found in the oil of this plant.
4 Describe the exact method you used to obtain oil from this plant.
5 Describe any qualitative observations you had about the oil obtained (e.g. smell, colour and viscosity).
6 Describe any quantitative observations you had about the oil you obtained (e.g. what volume of oil did you obtain from the mass of leaves used?)
7 List two safety precautions considered in this experiment.

12.3

PRACTICAL:
CLASSIFICATION & IDENTIFICATION

Can you identify chiral centres in popular medicinal compounds?

Practical worksheet
12.3 Can you identify chiral centres in popular medicinal compounds?

Context

Chiral compounds contain at least one chiral carbon atom. There are carbon atoms that are bonded to four different atoms or groups of atoms. Different configurations of these atoms or groups of atoms around the chiral carbon atom are called stereoisomers, and they are not superimposable upon their mirror image. One example is bromochlorofluoromethane (CHBrClF).

(*R*)-bromochlorofluoromethane **(*S*)-bromochlorofluoromethane**

FIGURE 1 There are two stereoisomers of bromochlorofluoromethane due to its chiral carbon atom.

Many biological compounds are chiral. This means that many medicinal compounds bind to targets in the body in a stereospecific way.

In this practical, you will analyse the top 10 small molecule medicines in Australia (2020) and identify the chiral carbons.

Aim

To identify chiral carbon atoms in the structural formulas of the medicinal compounds presented

Questions

1 Define the term *stereoisomer*.

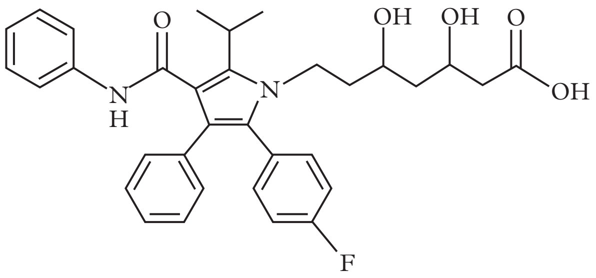
2 Define the term *chiral carbon atom*.

3 Draw the structural formulas for the two stereoisomers of 2-chlorobutane.

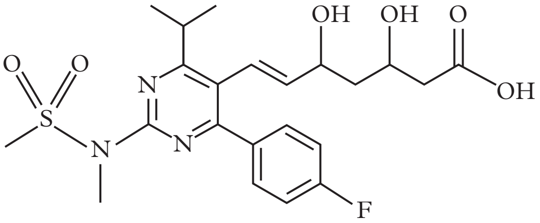
4 Draw the structural formula of one more small organic compound with a chiral carbon atom and write its name. **Note:** There is no need to add the (*R*) or (*S*) notation.

5 Identify the chiral carbon atoms in each of the structural formulas of medicinal compounds presented below.

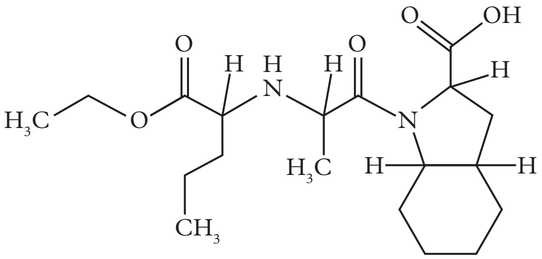
Atorvastatin



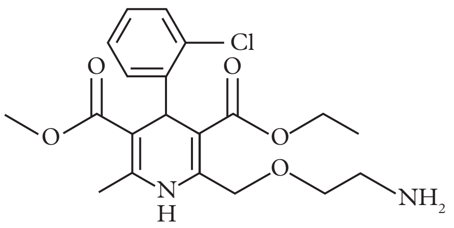
Rosuvastatin



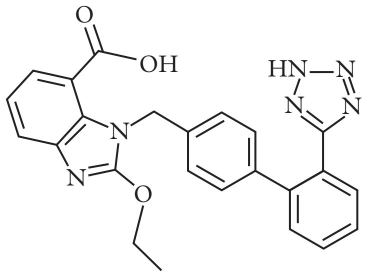
Perindopril



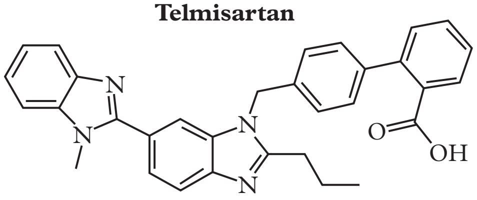
Amlodipine



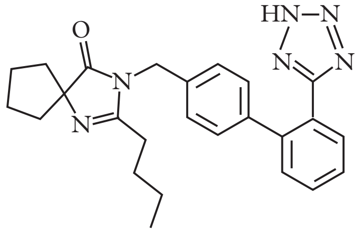
Candesartan



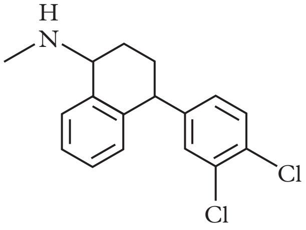
Telmisartan



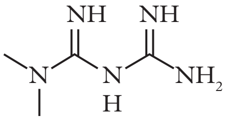
Irbesartan



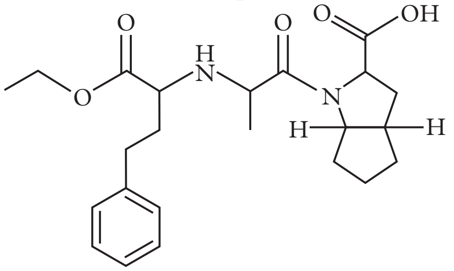
Sertraline



Metformin



Ramipril



6 The top two medicines belong to the statins group. They are used to lower cholesterol. Compare the structures of these two medicines. Redraw them in your logbook and highlight the functional groups that you predict would be important for the activity of the medicines.

7 The sartans (compounds 5–7) are used to lower blood pressure. Evaluate the structures of these three medicines. Redraw them in your logbook and highlight the functional groups that you predict would be important for the activity of the medicines.

8 Analyse the structures of compounds 3, 4, 8, 9 and 10. Identify any similarities in their structure. Based on their functional groups and structures alone, predict whether they may bind to similar targets to those of the statins and sartans.

9 Conduct some research on the other compounds and find out what they are used to treat.

   a Describe what you can conclude about the most common diseases in Australia, based on this list of medicines.

   b Conduct some research to determine whether you were right about your answer in Question 8.

12.4
PRACTICAL: MODELLING

How can paper be used to model a simple protein?

Video demonstration
12.4 How can paper be used to model a simple protein?

Practical worksheet
12.4 How can paper be used to model a simple protein?

Risk assessment
12.4 How can paper be used to model a simple protein?

Context

Proteins are formed in ribosomes by condensation polymerisation reactions. Almost immediately after this, they fold into specific three-dimensional shapes by forming bonds between the side groups of the amino acid residues in the chain. The specific three-dimensional shape, called the tertiary structure, is determined by the sequence of amino acids in the chain.

There are four types of bonds formed between side chains:

- Covalent bonds between cysteine amino acid residues
- Hydrogen bonds between polar amino acid residues
- Ionic interactions between acidic and basic amino acid residues
- Dispersion forces between non-polar residues

In this practical, you will use paper to model the 3D structure of a protein.

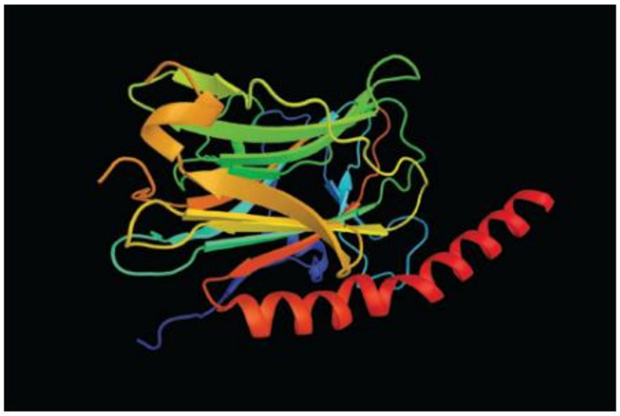


FIGURE 1 Calsequestrin protein (found in human muscle cells) folded into a specific three-dimensional shape

Aim

To model a protein using a long piece of paper by connecting amino acid side chains in a realistic way

Materials

- Two A3 sheets of paper (42 cm × 29 cm)
- Five coloured pens (preferably red, blue, green, yellow and black)
- 10 paperclips

Method

1 Make a random sequence of 10 large dots spaced evenly along the length of your strip of paper. These dots represent the five different types of amino acid you will model in your protein. Try to use the same number of each type (i.e. two of each). Record the colours you used in Table 1.

TABLE 1 Colour guide for amino acids

Amino acid	Type of side chain	Preferred colour	Colour used
Glutamic acid	Acidic	Red	
Arginine	Basic	Blue	
Cysteine	Sulfur-containing	Yellow	
Serine	Polar	Green	
Leucine	Non-polar	Black	

FIGURE 2 An example of how your strip of paper could look. The exact sequence of amino acids will be different each time you do this activity.

2 Photograph your strip of paper or redraw it in your logbook. You could also write the sequence of colours used.

3 Connect the amino acids according to how they could connect in a real protein. Start at one end of the paper and connect the following dots using paperclips:
    - If the dot is red (acidic side-chain), connect it to the next blue dot (basic side-chain).
    - If the dot is blue (basic side-chain), connect it to the next red dot (acidic side-chain).
    - If the dot is yellow (sulfur-containing side-chain), connect it to the next yellow dot.
    - If the dot is green (polar side-chain), connect it to the next green dot.
    - If the dot is black (non-polar side-chain), connect it to the next black dot.

4 Your protein has now been folded into a specific three-dimensional shape.

Questions

1 Write the primary structure (amino acid sequence) of the protein you created using the correct three-letter codes. These can be found in your data book.

2 Explain what might happen to the structure if:
    a temperature is significantly increased
    b temperature is significantly decreased
    c the protein is placed in an alkaline solution ($pH > 10$)
    d the protein is placed in an acidic solution ($pH < 4$).

3 Make another protein using the same sequence of dots but with two dots in swapped positions. This represents a mutation in the protein.
    a Compare the three-dimensional shapes of the two proteins you have folded.
    b Suggest why a mutation in a protein may make it unable to bind to its target molecule. Refer to the three-dimensional structure of the active site in your answer.

12.5
PRACTICAL: LITERATURE REVIEW

Does the lock-and-key model still fit?

Practical worksheet
12.5 Does the lock-and-key model still fit?

Context

Medicines often function as competitive enzyme inhibitors. These compounds have historically been thought to bind to enzymes through a lock-and-key mechanism. However, there are competing theories called the induced fit and transition-state models.

In this practical, you will compare the lock-and-key model with alternative models of enzyme binding.

FIGURE 1 Enzyme binding is often described using the lock-and-key model.

Aim

To evaluate the use of the lock-and-key model and compare this to alternative models

Instructions

1 Look for secondary sources of information. These can include the internet, books, scientific magazines, videos, podcasts, or interviews with experts.

2 Use the CRAAP method to evaluate the reliability of your sources.

3 Make notes about what you have learnt. You can organise your information in different ways.

4 Prepare a 4–5-minute oral presentation video that answers the following questions:

  a How does the lock-and-key mechanism work?

  b How are the induced-fit and transition-state models different from the lock-and-key model?

  c What factors affect binding of the substrate or medicine to the enzyme? Do they differ depending on the model used? If so, describe these differences.

  d Is one model more correct than the others for describing the interaction between a substrate or medicine with an enzyme? Justify your response.

  e Explain why different theories and models exist to explain the same scientific concept.

5 Make sure you record the details of all of the sources you use, including the title of the source, who it was written by, when it was written, page numbers or URLs and the date you accessed it.

Refer back to Chapter 13 for specific guidance on completing your literature review.

ANSWERS

You can find the full worked solutions to the answers in your **Student obook pro**

Chapter 1: Chemistry toolkit

GROUNDWORK

1A A hypothesis is a testable statement that predicts how the independent variable will affect the dependent variable in an investigation.

1B Qualitative data is data described in words, phrases or categories (e.g. colour of the solution). Quantitative data is data described using numbers, quantities or other numerical values (e.g. concentration of NaCl).

1C Student answers will vary, but examples include: title, units, columns, *x*- and *y*-axes, independent variable, dependent variable.

1D A controlled variable is one that is kept the same during an experiment. An independent variable is what you change – it is the variable being investigated. A dependent variable is what you measure in an experiment.

1.1 Overview of VCE Chemistry

1.1 CHECK YOUR LEARNING

1 Areas of Study 1 and 2: one of the assessment tasks outlined in the list (e.g. a report of laboratory or fieldwork activity) Area of Study 3: a report of a student-designed and student-conducted scientific investigation using a structured scientific poster and logbook entries.

2 Student answers will vary.

3 **a–e** Student answers will vary.

4 Student answers will vary.

1.2 Aboriginal and Torres Strait Islander knowledge, cultures and histories

1.2 CHECK YOUR LEARNING

1 Student answers will vary.

2 Aboriginal and Torres Strait Islander Peoples have developed and refined their own chemical knowledge over thousands of years; it is highly sophisticated and specialised for the specific location where they live.

3 Student answers will vary but should include how plants are used as medicine, and how natural materials are used and modified for a particular purpose.

4 Student answers will vary.

1.3 Developing aims, questions and hypotheses

1.3 CHECK YOUR LEARNING

1 The scientific process is a cycle in which the results are used as feedback to perform additional tests or make new hypotheses.

2 Aim: a statement of what is to be investigated. Hypothesis: a testable statement that predicts how the independent variable affects the dependent variable in the investigation, including an explanation.

3 **a** Student answers will vary but should include a clear, concise statement about what they are investigating and provide enough scope for experimentation.

b Student answers will vary but should include a question of what is to be investigated – e.g. ‘What effect do acids and bases have on the growth of plants?’

c Student answers will vary but should include a testable statement that predicts how the independent variable will affect the dependent variable and include a scientific explanation for the prediction – e.g. ‘If acids are used on plants in the garden, then their growth will be reduced because a lower soil pH limits the availability of essential nutrients for the plant.’

1.4 Planning and conducting investigations

1.4 CHECK YOUR LEARNING

1 A logbook is a useful record of primary and secondary data for scientific investigations.

2 Primary data is raw data that is collected directly from sources, whereas secondary data is data that has already been collected and is then analysed and interpreted.

3 **a** Similarity: Both can involve testing an independent variable to observe its effects on the dependent variable. Difference: Controlled experiments are typically conducted in a lab, whereas fieldwork is undertaken at a specific location.

b Similarity: Both involve looking at information to solve a problem. Difference: Case studies involve solving problems in real or hypothetical situations using knowledge you already have or information that has been given to you, whereas literature reviews require you to look for secondary data to answer a question.

c Similarity: Both involve investigation of a real system or scientific phenomena using visualisation tools. They can both be used to make predictions. Difference: Modelling requires you to create a model, whereas simulation involves using existing models to understand behaviour of systems.

4 Student answers will vary but should include a written question that clearly explains the chemical issue being investigated and the experimental approach being used.

1.5 Safety in chemistry

1.5 CHECK YOUR LEARNING

1 Chemists often work with dangerous goods and hazardous substances; therefore, it is important to follow safety rules and conduct risk assessments for experimental investigations in the laboratory.

2 Student answers will vary but should include an evaluation of each section of the risk assessment before making a final risk judgment of the laboratory-based activity.

3 Similarities: potential hazards and standard handling procedures are listed. Differences: hazard warning symbols and a classification of the type of hazardous substance are provided for GHS substances classified as hazardous.

4 Student answers will vary but should include who has prepared the risk assessment, the title of the investigation, who is conducting the investigation, the date and location, a list of materials, procedure, equipment and chemicals used, and a final risk judgment with a signature.

1.6 Ethical understanding

1.6 CHALLENGE

1 Student answers will vary.

2 Student answers will vary but should identify issues such as the impact or harm of the issue (on living things or the environment), integrity, safety and the needs of society in the present or future.

3 Student answers will vary but could include political, legal, economic or social factors.

4–7 Student answers will vary.

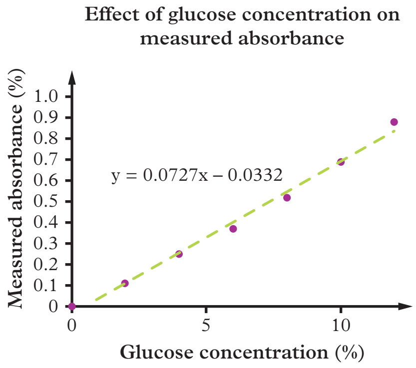
1.6 CHECK YOUR LEARNING

1. To make sure the broader implications of conducting the investigation are considered, reporting of results is honest and all contributors to the ideas are acknowledged.
2. Student answers will vary but should include a discussion of the similarities and differences.
3. Student answers will vary but should include an explanation of arguments for and against the ethical issue, as well as justification for opinions made or decisions taken on the issue.

1.7 Generating, collating and recording data

1.7 CHECK YOUR LEARNING

1. A title, labelled axes, the IV plotted on the *x*-axis, the DV plotted on the *y*-axis, appropriate axis scales, a key (if applicable)
2. Column graphs display categorical data, with bars that are plotted vertically or horizontally but are not touching. Histograms show the frequency of numerical data as a percentage; each column width can be a single data value or a data interval.
3. **a** The distilled water is used as a controlled variable.
   b DV: Measured absorbance
   IV: Glucose concentration
   c Student answers will vary but should present the data using a line graph.


Effect of glucose concentration on measured absorbance

Students should note that as glucose concentration increases, the measured absorbance increases.

1.8 Evaluating data and investigations

1.8 CHECK YOUR LEARNING

1. Systematic errors (caused by a problem with the method or equipment)
2. It reduces the impact of random errors in the measurement process. By conducting additional trials, more data can be obtained and therefore the data can be averaged.
3. **a** Accuracy: how close a value is to the true value. Precision: a measure of how close the data values in a set are to each other.
   b Repeatability: the data values can be produced again (by the same experiment and under the same conditions). Reproducibility: the same data values can be produced again (under slightly different conditions).
   c Mistakes: personal errors that should not be included in reporting or data analysis. Errors: can be random (unpredictable errors that occur because of an error in the measurement process) or systematic (consistent and repeatable errors that are caused by a problem with the method or equipment).
4. The data obtained is not accurate because the values are not close to the accepted boiling point of water. The values in the data set lack precision.
5. **a** 4.08 **b** 36.0 **c** 2.90
6. Student answers will vary.

1.9 Constructing evidence-based arguments and conclusions

1.9 CHECK YOUR LEARNING

1. Evidence-based arguments are important to properly communicate scientific ideas that are backed by evidence. They are primarily used in the analysis, discussion and conclusion sections of the scientific investigation report.
2. Three of: use scientific terminology and/or supporting figures, make evaluative statements, use formal language and avoid first person.
3. Student answers will vary but should include an analysis of the experimental results, errors, inconsistencies and outliers that may have affected the results, scientific explanations that link the results to the question and aim, and suggestions for improvement.

1.10 Communicating

1.10 CHECK YOUR LEARNING

1. **a** Identify and understand your target audience, use appropriate language, science literacy and convention, essential content and stylistic elements to enhance presentation.
   b Using stylistic elements (e.g. diagrams, photos, graphs, tables) so that the content is clearly conveyed, easily understood and enhances written content
2. Scientific language and literacy should be used throughout a poster so that a summary of the research is communicated clearly with the intended audience.
3. Student answers will vary but should include a discussion of the similarities and differences.
4. The results present the raw data collected and recorded during the investigation in the most appropriate format. The discussion examines whether the data supports or refutes the hypothesis, analyses the data and compares to expected results, identifies ambiguities and further questions that may arise, and provides explanations of issues with the data.
5. Student answers will vary but should include important information and be written for the intended audience.

1.11 Sustainability

1.11 CHECK YOUR LEARNING

1. Atom economy, catalysis, design for degradation, design for energy efficiency, design of safer chemicals, prevention and use of renewable feedstocks.
2. A linear economy removes resources from the ground to make products that are then disposed of, whereas a circular economy sustainably focuses on using and reusing resources repeatedly at different stages of production and consumption.
3. Student answers will vary but should include a discussion about which of the 12 principles of green chemistry were important when considering replacing the hazardous substance.

1.12 Using the periodic table

1.12 CHECK YOUR LEARNING

1. Electronegativity of non-metals increases going across a period and up a group on the periodic table.
2. The metallic character of transition metals decreases going across a period and up a group on the periodic table.

3

Atom	Name of element	Number of protons	Number of neutrons
$^{32}_{15}P$	Phosphorus	15	17
$^{238}_{92}U$	Uranium	92	146
$^{63}_{31}Ga$	Gallium	31	32

4 Caesium is at the bottom left of the periodic table and fluorine is on the top right. As reactivity increases going down a group and decreases going across a period, we would expect caesium to be more reactive than fluorine.

1.13 Balancing chemical equations

1.13 CHECK YOUR LEARNING

1 The law of conservation of mass states that in an isolated system, during a chemical reaction, mass is conserved; atoms cannot be created or destroyed. The mass of reactants and products remains constant, and the number of atoms present does not change.

2 **a** $2Na(s)+Cl_2(g)\rightarrow 2NaCl(s)$

b $4Fe(s)+3O_2(g)\rightarrow 2Fe_2O_3(s)$

c $Ag_2CO_3(s)\rightarrow Ag_2O(s)+CO_2(g)$

d $C_2H_4(g)+2O_2(g)\rightarrow 2CO(g)+2H_2O(g)$

e $P_4(s)+O_2(g)\rightarrow 2P_2O_5(s)$

f $Fe_2O_3(s)+3CO(g)\rightarrow 2Fe(s)+3CO_2(g)$

g $2HCl(aq)+Mg(s)\rightarrow MgCl_2(aq)+H_2(g)$

h $Ca(CN)_2(s)+2HCl(aq)\rightarrow 2HCN(aq)+CaCl_2(aq)$

i $2Fe(s)+3Cl_2(g)\rightarrow 2FeCl_3(s)$

j $2Mg(s)+O_2(g)\rightarrow 2MgO(s)$

3 Student answers will vary.

1.14 Preparing for assessment

1.14 CHECK YOUR LEARNING

1 Two of: use writing time effectively, note down ideas once writing time has commenced, use writing space provided or ask for spare paper, don't dwell on a difficult question but continue with other questions and return to the question later, eliminate incorrect answers for multiple-choice questions, beware of careless mistakes and draw large, clear diagrams when required.

2 It is important to have a work–life balance; to have time to study, participate in extra-curricular activities and rest.

3 Student answers will vary but should include a goal that is clearly defined, can be measured, and is achievable, worthwhile and time-based.

4 **a** *Calculate* asks students to find a numerical answer, whereas *deduce* asks them to use the information to provide a new relationship.

b *Outline* asks students to provide an overview of an argument, text or an image, whereas *identify* only requires them to determine a characteristic.

c *Evaluate* asks students to calculate a value or make a judgment, whereas *justify* asks them to provide reasoning for an opinion/argument.

5 Student answers will vary.

Chapter 1 review

MULTIPLE CHOICE

1	D	**2**	B	**3**	C	**4**	C
5	B	**6**	D	**7**	B	**8**	A

SHORT ANSWER

9 Student answers will vary.

10 Use connectives to show cause and effect, scientific terminology and/or supporting figures, supporting facts and/or quotes from experts or other researched sources, make evaluative statements, use formal language and avoid third person.

11 **a** DV: voltage (V)
IV: temperature (°C)

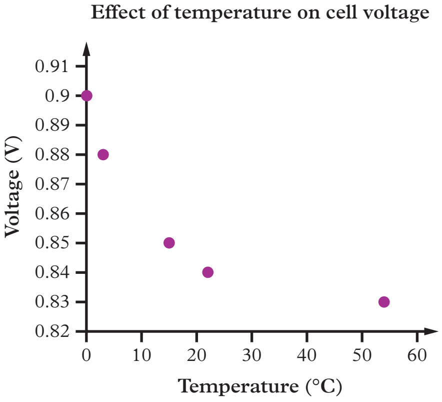
b To determine the effect of temperature on electrochemical cell voltage

c If temperature increases, then the voltage of the cell will change.

d

Temperature °C	Average voltage
0	0.90
3	0.88
15	0.85
22	0.84
54	0.83

e Student answers will vary but should present the data using a line graph.



Students should note that as temperature increases, the voltage decreases. Students may also note potential outliers at 0°C or 3°C.

f 77% (2 sig fig)

12 Student answers will vary but should include an analysis of the experimental results, errors, inconsistencies and outliers that may have affected the results, scientific explanations that link the results to the question and aim, and suggestions for improvement.

13 Student answers will vary but should include who has prepared the risk assessment, the title of the experimental investigation, information on who is conducting the investigation as well as the date and location, a list of materials to be prepared by the laboratory technician, a record of the procedure as well as all equipment and chemicals used, and a final risk judgment, with a signature included.

Chapter 2: Carbon-based fuels

GROUNDWORK

2A Heat, chemical and wind

2B Releasing heat into the surrounding environment

2C Fossil fuels are fuels that are found on Earth, which have been created through decomposing plants and animals millions of years ago.

2.1 The renewability of fuels

2.1 REAL-WORLD CHEMISTRY

1 Bioethanol is a renewable source, meaning that it can be replenished in a relatively short period of time. It also produces lower carbon dioxide gas emissions than petrol, so it has a lesser impact on global warming.

2 Other countries do not have the same supply of sugar cane, which is needed to produce bioethanol. They also might not have the machinery/technology to synthesise bioethanol.

3 Brazil would be left without fuel. The cost of producing the bioethanol would be high.

2.1 CHECK YOUR LEARNING

1 A fuel is a substance that provides heat energy when it is reacted with another substance.

2 Biogas is produced by the anaerobic fermentation of vegetable and animal waste by bacteria.

3 Biodiesel: made from renewable resources such as plant materials, which can always be replenished in a relatively short amount of time. Petrodiesel: relies on fossil fuels, which take longer to be created and are in short supply.

4 Fossil fuels: take years to form underground as the decomposing plants and animals are compressed over millions of years by the pressure of the ocean above; non-renewable. Biofuels: made from plants that are readily available and quickly grown; renewable.

5 Biodiesel is renewable; petrodiesel is non-renewable. Biodiesel is more viscous, and has a slightly lower energy content and a higher melting point.

6 All three biofuels are renewable, as they are sourced from plants that are fast-growing and high-yielding species.

7 The processes for producing biofuels are quite expensive as they require specific machinery and reactants to produce the biofuel.

2.2 Fuel sources for the body

2.2 REAL-WORLD CHEMISTRY

1 High fructose corn syrup and cane sugar

2 Aspartame, acesulfame potassium (ace-K), stevia leaf, sucralose

2.2 CHALLENGE

7.9×10^2 kJ (2 sig fig)

2.2 CHECK YOUR LEARNING

1 **a** $16\,\text{kJ}\,\text{g}^{-1}$
  b $17\,\text{kJ}\,\text{g}^{-1}$
  c $37\,\text{kJ}\,\text{g}^{-1}$

2 The energy released on combustion is the amount of energy released when a macronutrient is combusted, whereas energy content is the amount of energy a macronutrient provides that can be used by the body.

3 $C_{18}H_{32}O_2 + 25O_2 \rightarrow 18CO_2 + 16H_2O$

4 25.5 g (3 sig fig)

5 7.3×10^2 kJ (2 sig fig)

6 Different chemical compositions (i.e. different lengths of hydrocarbon chains or presence and number of carbon–carbon double bonds).

2.3 Photosynthesis

2.3 SKILL DRILL

1 There would be a plentiful supply of renewable sources that grow faster or produce more, that can be manufactured to become biofuels.

2 Having enough sunlight and carbon dioxide for the reaction to occur

3 Student answers will vary, but would have to consider whether the cost outweighs the advantages.

2.3 CHECK YOUR LEARNING

1 $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

2 Light intensity, carbon dioxide concentration, optimal temperature

3 Endothermic reaction

4 grass < rice < pineapple

5 The chemical composition of food crops is different from that of non-food plants. Food crops are likely to utilise the sunlight better when carrying out photosynthesis.

2.4 Oxidation of glucose

2.4 REAL-WORLD CHEMISTRY

1 Made from organic compounds; lightweight

2 Increase in bowel movements as the person would be eating more starch

2.4 CHECK YOUR LEARNING

1 Chemical energy

2 $C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$

3 Glucose, fructose and galactose

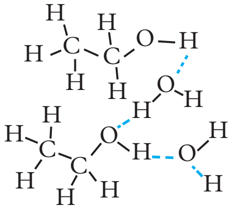
4 Amylose: straight-chain polymer of D-glucose units; Amylopectin: branched-chain polymer of D-glucose units

5 The straight chains of amylose make it harder to break down and digest (great number of intermolecular forces). Branched amylopectin chains are unable to pack together as tightly and have fewer intermolecular forces, making them easier to break apart.

2.5 Producing bioethanol

2.5 CHALLENGE

1 Ethanol is able to dissolve in water due to the presence of the hydroxyl group, which allows it to form hydrogen bonds with water.



2 Ethanol is able to dissolve in petrol due to the presence of the non-polar alkane chain, which can form dispersion forces with the non-polar molecules in petrol.

3 Water, as the hydrogen bond between the ethanol molecule and water would form more readily and is a stronger intermolecular bond than dispersion forces

2.5 CHECK YOUR LEARNING

1 E10 fuel is a mixture of 90% petrol and 10% bioethanol.

2 The temperature should be warm enough and the region must have enough sunlight.

3 $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$

4 Cellulose, hemicellulose and lignin

5 Using enzymes

6 Ethanol is distilled to purify the fuel. This is done by separating ethanol and water using their different boiling points. When the mixture is heated slightly above 78°C, the ethanol will boil, but the water will not. The ethanol gas formed can be condensed and collected in a separate container at its highest purity.

2.6 Comparing exothermic and endothermic reactions

2.6 CHECK YOUR LEARNING

1 Endothermic: vaporisation, melting, sublimation, boiling
  Exothermic: freezing, condensation, deposition

2 **a** Positive
  b Negative

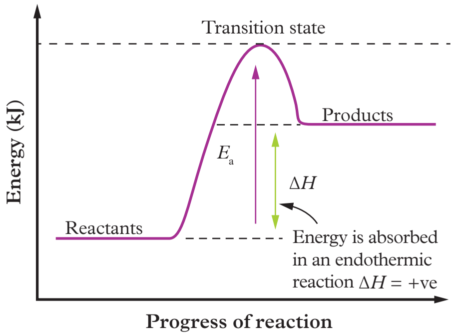
3 absorbed; released

4 $\Delta H_{\text{endothermic}} > 0$
  $\Delta H_{\text{exothermic}} < 0$

5 Reactants

6 The second chemical equation (where water vapour is produced) would have the higher ΔH value, as more energy is required to break the bonds to form a gaseous product.

7



8 $E_a(\text{respiration}) = E_a(\text{photosynthesis}) - \Delta H(\text{photosynthesis})$

2.7 Limiting reactants and reagents

2.7 REAL-WORLD CHEMISTRY

1 $2Al(s)+Fe_2O_3(aq)\rightarrow Al_2O_3(aq)+2Fe(s)$

2 0.91 kg (2 sig fig)

2.7 CHECK YOUR LEARNING

1 The limiting reactant is the substance that, when depleted, results in the end of a reaction.

2 Substance A

3 1.7 mol ≈ 2 mol (1 sig fig)

4 **a** $2Mg(s)+O_2(g)\rightarrow 2MgO(s)$
  b Mg
  c 0.12 mol (2 sig fig)

5 **a** $C_3H_8(g)+5O_2(g)\rightarrow 3CO_2(g)+4H_2O(g)$
  b O_2
  c 0.95 mol (2 sig fig)

6 49 mL (2 sig fig)

2.8 Combustion of fuels

2.8 CHECK YOUR LEARNING

1 Carbon monoxide, soot and water

2 Heat of combustion is the energy released when a specific amount of fuel is burnt during a complete combustion reaction. Enthalpy of combustion is the change in reaction energy that occurs due to complete combustion.

3 Both complete and incomplete combustion require oxygen as a reactant. Incomplete combustion produces carbon monoxide and water, whereas complete combustion only produces carbon dioxide and water.

4 **a** $2C_8H_{18}+25O_2\rightarrow 16CO_2+18H_2O$
  b $2C_4H_{10}+9O_2\rightarrow 8CO+10H_2O$
  c $C_6H_{12}O_6+6O_2\rightarrow 6CO_2+6H_2O$
  d $2C_2H_2+3O_2\rightarrow 4C+2H_2O$

5 **a** $C_2H_5OH(l)+3O_2(g)\rightarrow 2CO_2(g)+3H_2O(l) \quad \Delta H=-1360\,kJ\,mol^{-1}$
  b $2CO_2(g)+3H_2O(l)\rightarrow C_2H_5OH(l)+3O_2(g) \quad \Delta H=1360\,kJ\,mol^{-1}$

6 When the thermochemical equation is reversed, the magnitude of ΔH will remain the same; the signs are reversed as it changes from an exothermic to an endothermic reaction.

7 **a** 9990 kJ or 9.99×10^3 kJ (3 sig fig)
  b 353 kJ (3 sig fig)

8 Student answers will vary, but should consider the release of carbon monoxide as a greenhouse gas, contributing to global warming.

Chapter 2 review

MULTIPLE CHOICE

1 A **2** D **3** A **4** A **5** D
6 D **7** A **8** D **9** D **10** A

SHORT ANSWER

11 $C_6H_{12}O_6(aq)+6O_2(g)\rightarrow 6CO_2(g)+6H_2O(l)$

12 $6CO_2(g)+6H_2O(l)\xrightarrow{sunlight} C_6H_{12}O_6(aq)+6O_2(g)$

13 **a** Starch 2
  b Starch 1
  c Amylopectin is more soluble in water as the branches can easily be hydrolysed.

14 A fuel is a compound that can be reacted to produce energy.

15 The oxidation of carbohydrates, proteins and lipids rely on different enzymes. They have varying abilities to oxidise and extract nutrients from these three different fuels.

16 Use of renewable feedstocks: plants used to make E10 are available in a more ready supply than fossil fuels, from which petrol is derived.

17 $C_6H_{12}O_6\,(aq)\xrightarrow{enzymes} 2C_2H_5OH(aq)+2CO_2(aq)$

18 Fast-growing, efficiently photosynthetic plants, such as corn, are grown and harvested. Glucose is then extracted from the plants through mechanical and chemical processing. The glucose is then fermented (reacted with yeast enzymes in the absence of oxygen) to produce ethanol and carbon dioxide gas. The ethanol is distilled to increase its purity. Finally, the ethanol is added to petrol at a concentration of 10%.

19 Bioethanol is renewable, meaning that it can be replenished relatively quickly compared to petrodiesel. It also decreases the dependence on fossil fuels as it is produced locally and in small batches. Finally, it has lower carbon emission.

20 Bioethanol is produced from plants, which need to be grown relatively quickly. Creating a suitable environment to grow the crops can be a challenge if the temperature is not right. Also, the land required to grow the crops for fuel production could be used to grow food.

21 **a** –1568 kJ
  b –3940 kJ
  c Exothermic

22 Melting, boiling and sublimation

23 **a** $C_2H_5OH(aq)+3O_2(g)\rightarrow 2CO_2(g)+3H_2O(l)$
  b Enthalpy of reactants > enthalpy of products
  c 29600 kJ or 2.96×10^4 kJ (3 sig fig)

24 **a** $C_3H_8(g)+5O_2(g)\rightarrow 3CO_2(g)+4H_2O(g)$
  b 2522 kJ or 2.52×10^3 kJ (3 sig fig)

25 The blue flame is the heating flame, whereas the yellow flame is the safety flame. When the Bunsen burner air hole is open, air can flow into the system and mix with the natural gas to ensure complete combustion with a blue flame and a high temperature. When the hole is closed, then incomplete combustion will occur, leading to the yellow flame and the production of tiny particles of solid carbon called soot. It will also have a lower temperature.

26 **a** $C_9H_{20}(g)+14O_2(g)\rightarrow 9CO_2(g)+10H_2O(g)$
  b Reactants
  c Products

27 **a** A
  b 0.70 mol (2 sig fig)

28 **a** X
  b 0.020 mol (2 sig fig)

29 **a** $C_2H_2(g)+3O_2(g)\rightarrow 2H_2O(l)+2CO_2(g)$
  b O_2
  c 1.2 g (2 sig fig)

30 HCl

Chapter 3: Measuring changes in chemical reactions

GROUNDWORK

3A The amount of chemical energy stored within the bonds of a substance is called its enthalpy, or heat content, and is represented by the symbol *H*.

3B To combust a fuel, it must be burned in the presence of oxygen. This reaction releases heat energy, carbon dioxide and water vapour. If combustion is incomplete, carbon monoxide can be formed.

3C Specific heat capacity is a measure of the amount of heat energy, in joules, that it takes to increase the temperature of a specific volume of water by 1°C.

3.1 Stoichiometry involving the combustion of fuels

3.1 CHALLENGE

1 192 mol or 1.9×10^2 mol (2 sig fig)

2 384.6 mol or 3.8×10^2 mol (2 sig fig)

3 14397.3 L or 1.4×10^4 L (2 sig fig)

3.1 CHECK YOUR LEARNING

1 Liquid

2 Carbon dioxide; greenhouse gases can trap infrared radiation and prevent it from leaving the atmosphere. This causes the temperature of the Earth's atmosphere to rise, which affects other biomes.

3 **a** $2C_3H_6(g)+9O_2(g)\rightarrow 6CO_2(g)+6H_2O(l)$

b $4Al(s)+3O_2(g)\rightarrow 2Al_2O_3(s)$

c $N_2(g)+O_2(g)\rightarrow 2NO(g)$

d $2C_4H_6(g)+11O_2(g)\rightarrow 8CO_2(g)+6H_2O(l)$

e $C_6H_{12}O_6(aq)+6O_2(g)\rightarrow 6CO_2(g)+6H_2O(l)$

4 33688 kJ or 3.4×10^4 kJ (2 sig fig)

5 306 g or 3.1×10^2 g (2 sig fig)

6 Using $q = \Delta H\times n$: 1194 MJ or 1.2×10^3 MJ (2 sig fig)
Using $E = \Delta H\times m$: 1195 MJ or 1.2×10^3 MJ (2 sig fig)

7 Using $q = \Delta H\times n$: 8536 MJ or 8.5×10^3 MJ (2 sig fig)
Using $E = \Delta H\times m$: 8543 MJ or 8.5×10^3 MJ (2 sig fig)

8 101 L or 1.0×10^2 L (2 sig fig)

3.2 Specific heat capacity

3.2 CHECK YOUR LEARNING

1 Specific heat capacity is a measure of the amount of heat energy, in joules, that it takes to increase the temperature of 1 g of substance by 1°C.

2 If the mass of water is measured in g, and used in the calculation as g, the energy calculated will be in J. If the water is measured in kg, or converted from g to kg, and used in the calculation as kg, then the energy calculated will be in kJ.

3 A denser substance would need more energy to increase the temperature of 1 g of that substance by 1°C, so it will have a higher specific heat capacity.

4 9.3 kJ (2 sig fig)

5 37°C (2 sig fig)

6 5.4 g (2 sig fig)

7 Water in liquid form has particles that are closer together; in ice, the particles are in a rigid structure. Hence liquid water will have a higher specific heat capacity than ice due to its increased density, as it takes more energy to heat up the liquid water compared to the ice.

8 In liquid water, the particles are closer together and move around each other constantly, unlike in ice, so it has more space to move and absorb the heat being provided. Additionally, there is hydrogen bonding between water molecules, which requires a lot of energy to break.

3.3 The principles of solution calorimetry

3.3 CHALLENGE

9°C (1 sig fig)

3.3 SKILL DRILL

1 Random error due to different amounts of heat energy being lost due to the change in the outside temperature of the calorimeter

2 Insulation and ensuring the lid is tightly fitted

3 **a** Lower because of a degradation of the insulated container; higher calibration factor

b Validity: the results are still answering the aim of the experiment. Accuracy: the results are not accurate, because the incorrect calibration factor was used.

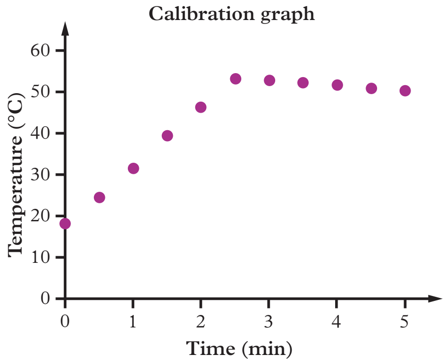
3.3 CHECK YOUR LEARNING

1 A calorimeter can measure the temperature change of the solution placed in it which can be used to calculate the enthalpy of the solution.

2 A calorimeter is insulated so that it can minimise the heat loss to the surrounding environment. If it was not insulated, it would take longer for the solution to heat up.

3 Combustion of solids, liquids and gases requires an ignition source and excess supply of oxygen, which a solution calorimeter does not accommodate.

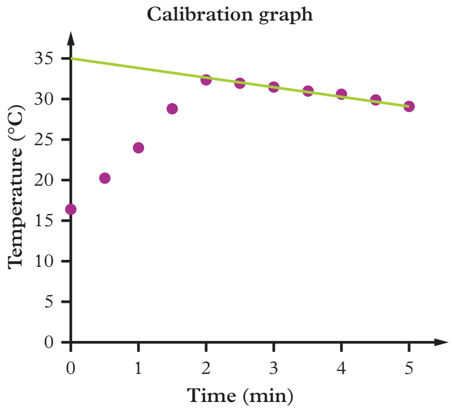
4 **a**



b 37.9 °C

c 50 J °C⁻¹ (2 sig fig)

5 **a**



b 18.5 °C

c 39 J °C⁻¹ (2 sig fig)

6 +34 kJ mol⁻¹ (2 sig fig); endothermic.

7 20 kJ mol⁻¹ (2 sig fig)

8 13 kJ mol⁻¹ (2 sig fig)

9 It determines the amount of energy that it takes to change the temperature of the solution in the calorimeter by 1°C.

10 This is a systematic error. Because the student did not extrapolate the graph, their calibration factor is off, so the measured enthalpy of each reaction is incorrect.

3.4 Energy from fuels and food

3.4 REAL-WORLD CHEMISTRY

1 chemical calories = nutritional calories × 4.2

2 Men: 11 kJ (2 sig fig)
Women: 8.4 kJ (2 sig fig)

3.4 SKILL DRILL

1 50.3 kJ (3 sig fig

2 26.8 kJ (3 sig fig)

3 53.2% (3 sig fig)

4 The experimental value is significantly lower than the theoretical value, because heat would be lost to the surroundings and apparatus, which has not been accounted for.

5 This is a random error, because different amounts of heat could be lost depending on the environment and equipment used.

6 This experiment was valid because the results have enabled the student to measure the heat of combustion of ethanol.

3.4 CHECK YOUR LEARNING

1 To measure the heat of combustion of a fuel, we need to ignite it, which cannot be done in a solution calorimeter.

2 Fuels are often impure mixtures, which means that the energy from the fuel cannot be measured in kJ mol⁻¹.

3 Due to the massive amount of heat lost to the surroundings and apparatus, the experimental data may not be accurate. However, since relatively the same amount of heat has been lost, the experimental data would be similar or the same, so the results may be considered precise.

4 **a** 1730 kJ mol⁻¹ or 1.7×10^3 kJ mol⁻¹ (2 sig fig)

b 40% (2 sig fig)

5 **a** 40.1 kJ mol⁻¹ (3 sig fig)

b 10.8% (3 sig fig)

6 **a** fat: 6.8 g (2 sig fig)
carbohydrates: 0.15 g (2 sig fig)
protein: 8.2 g (2 sig fig)

b fat: 2.5×10^2 kJ (2 sig fig)
carbohydrates: 2.4 kJ (2 sig fig)
protein: 1.4×10^2 kJ (2 sig fig)
c 393 kJ or 3.9×10^2 kJ (2 sig fig)
d 4.5% (2 sig fig)

7 **a** fat: 2.5 g (2 sig fig)
carbohydrates: 12 g (2 sig fig)
protein: 8.8 g (2 sig fig)
b fat: 93 kJ (2 sig fig)
carbohydrates: 1.8×10^2 kJ (2 sig fig)
protein: 1.5×10^2 kJ (2 sig fig)
c 425.25 kJ or 4.3×10^2 kJ (2 sig fig)
d 4.9% (2 sig fig)

8 (1) the separation of the burning wick from the base of beaker (2) heat loss to the surrounding environment from the beaker or metal can and also from the wick (3) heat loss from the top of the beaker through water vapour, if a lid is not used.

9 Milk will provide more energy. Carbohydrates are digested and absorbed at a faster rate.

Chapter 3 review

MULTIPLE CHOICE

1	A	**2**	D	**3**	C	**4**	D	**5**	D
6	B	**7**	D	**8**	C	**9**	D	**10**	A

SHORT ANSWER

11 Specific heat capacity is a measure of the amount of heat energy, in joules, that it takes to increase the temperature of a specific volume of water by 1°C.

12 A molecule with stronger intermolecular interactions than water would have a higher specific heat capacity because it can absorb more heat.

13 As the change in temperature in the uninsulated calorimeter would be lower than in the insulated calorimeter, its calibration factor will be higher than the calibration factor of the insulated calorimeter.

14 Moisture-free wood implies that there are limited amounts of water molecules in the firewood, so it would not have water molecules absorbing the heat being provided.

15 A solution calorimeter is insulated and loses less heat energy to the environment whereas a spirit burner loses a significant amount of heat energy when heating the water in the flask.

A solution calorimeter measures the enthalpy of aqueous reactions whereas a spirit burner measures the enthalpy of a combustion reaction.

16 5.5×10^3 kJ mol^{-1} (2 sig fig)

17 3.8×10^2 kJ mol^{-1} (2 sig fig)

18 126 MJ (3 sig fig)

19 **a** 1.0×10^3 kJ mol^{-1} (2 sig fig)
b 52% (2 sig fig)

20 **a** fat: 0.69 g (2 sig fig)
carbohydrates: 0.4 g (1 sig fig)
protein: 18 g (2 sig fig)
b fat: 26 kJ (2 sig fig)
carbohydrates: 6 kJ (1 sig fig)
protein: 3.0×10^2 kJ (2 sig fig)
c 332.5 kJ or 3.3×10^2 kJ (2 sig fig)
d 3.8% (2 sig fig)

21 **a** fat: 8.5 g (2 sig fig)
carbohydrates: 11 g (1 sig fig)
protein: 8.5 g (2 sig fig)
b fat: 3.1×10^2 kJ (2 sig fig)
carbohydrates: 1.8×10^2 kJ (1 sig fig)
protein: 1.4×10^2 kJ (2 sig fig)
c 635 kJ or 3.4×10^2 kJ (2 sig fig)
d 7.3% (2 sig fig)

22 When using a spirit burner, soot may be seen coating the bottom of the water beaker if combustion is incomplete.

To ensure complete combustion of the fuel, make sure that the flame is not partially enclosed or obstructed; this will allow the flame to have enough oxygen from the air to complete the combustion reaction.

23 Ethanol is partially oxidised and can only react with three oxygen molecules, resulting in a lower enthalpy value of 1360 kJ mol^{-1}. Ethane is not oxidised and can react with 7/2 (or 3.5) oxygen molecules, resulting in a higher enthalpy of 1560 kJ mol^{-1}.

Chapter 4: Primary galvanic cells and fuel cells

GROUNDWORK

4A A redox reaction is a pair of electron transfer reactions that occur simultaneously: reduction and oxidation. As one reactant loses one or more electrons (oxidation), the second reactant gains them to form new products (reduction).

4B An oxidising agent causes oxidation and is itself reduced in the redox reaction (gains electron(s)). A reducing agent causes reduction and is itself oxidised in the redox reaction (loses electron(s)).

4C Biomass is a renewable organic material that comes from plants and animals, and contains stored chemical energy from the Sun. It can be burned directly to produce heat or converted into renewable liquid and gaseous fuels through a variety of processes.

4.1 Redox reactions

4.1 CHECK YOUR LEARNING

1 The degree of oxidation of a chemical species indicates the number of electrons gained or lost by the species when it bonds in a compound.

2 Decreases

3 Reduction occurs when a chemical species gains one or more electrons. Oxidation occurs when a chemical species loses one or more electrons.

4 **a** O: 0
b N: +4; O: −2
c S: +6; O: −2
d C: 0; H: +1; O: −2
e S: +1; H: +1; O: −2
f H: +1; O: −1
g S: +1; H: −1
h P: +5; O: −2

5 **a** Oxidising agent: OCl^- – the oxidation number of Cl changes from +1 to −1; itself undergoes reduction

Reducing agent: $Fe(OH)_3$ – the oxidation number of Fe changes from +3 to +4; itself undergoes oxidation

b Oxidising agent: VO_4^{3-} – the oxidation number of V changes from +5 to +4; itself undergoes reduction

Reducing agent: Fe^{2+} – the oxidation number for Fe changes from +2 to +3; itself undergoes oxidation

c Oxidising agent: $Cr_2O_7^{2-}$ – the oxidation number of Cr changes from +6 to +3; itself undergoes reduction

Reducing agent: C_2H_5OH – the oxidation number for C changes from +2 to +4; itself undergoes oxidation

6 **a** $Fe(OH)_3(aq)/FeO_4^{2-}(aq)$ and $OCl^-(aq)/Cl^-(aq)$
b $Fe^{2+}(aq)/Fe^{3+}(aq)$ and $VO_4^{3-}(aq)/VO^{2+}(aq)$
c $Cr_2O_7^{2-}(aq)/Cr^{3+}(aq)$ and $C_2H_5OH(l)/CO_2(g)$

4.2 Half-equations and overall equations

4.2 CHALLENGE

$$2Cr_2O_7^{2-}(aq)+16H^+(aq) + 3CH_3CH_2CH_2CH_2OH(aq) \rightarrow 4Cr^{3+}(aq)+11H_2O(l)+ 3CH_3CH_2CH_2COOH(aq)$$

4.2 CHECK YOUR LEARNING

1 The chemical species loses one or more electrons; the electrons are located on the right of the reaction arrow.

2 The chemical species gains one or more electrons; the electrons are located on the left of the reaction arrow.

3 The number of electrons in the oxidation half-equation must be equal to the number of electrons in the reduction half-equation as there is a direct transfer of electrons from one substance to another.

4 **a** Oxidation: $Fe(s) \rightarrow Fe^{3+}(s) + 3e^-$
Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(s)$
Overall: $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$

b Oxidation: $S(s) \rightarrow S^{6+}(g) + 6e^-$
Reduction: $F_2(g) + 2e^- \rightarrow 2F^-(g)$
Overall: $S(s) + 3F_2(g) \rightarrow SF_6(g)$

c Oxidation: $N_2(g) \rightarrow 2N^{2+}(g) + 4e^-$
Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(g)$
Overall: $N_2(g) + 2Cl_2(g) \rightarrow N_2Cl_4(g)$

d Reduction: $6CO_2(g) + 24H^+(aq) + 24e^- \rightarrow C_6H_{12}O_6(aq) + 6H_2O(l)$
Reduction: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$
Overall: $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

5 **a** Fe: reducing agent; Cl_2: oxidising agent

b S: reducing agent; F_2: oxidising agent

c N_2: reducing agent; Cl_2: oxidising agent

d H_2O: reducing agent; CO_2: oxidising agent

6 Student answers will vary but must include the key differences in the smelting of copper in ancient times compared to today.

7 Student answers will vary but must include the oxidation half-equation, reduction half-equation and overall equation.

4.3 Features of non-rechargeable galvanic cells

4.3 CHECK YOUR LEARNING

1 Galvanic cells consist of two half-cells in which a spontaneous redox reaction occurs. The spontaneous reaction takes place between the stronger reducing agent in the anode half-cell and the stronger oxidising agent in the cathode half-cell. The separate half-cells are connected by an outer circuit through which electrons flow from the anode to the cathode. The circuit is completed by an inner circuit consisting of a moist salt bridge where cations flow to the cathode and anions flow to the anode.

2 It completes the circuit and supplies ions into the half-cells to balance the charge.

3 **a** Cl_2 is the oxidising agent; Zn is the reducing agent.

b Zn is the anode; Pt is the cathode.

c $Zn(s) + Cl_2(g) \rightarrow Zn^{2+}(aq) + 2Cl^-(aq)$

4 **a** $H_2(g)/2H^+(aq)$; $Cl_2(g)/2Cl^-(aq)$

b $Cu(s)/Cu^{2+}(aq)$; $Fe(s)/Fe^{2+}(aq)$

c $Ag(s)/Ag^+(aq)$; $Pb(s)/Pb^{2+}(aq)$

5 **a** Zinc solid only

b Copper or inert electrode (graphite, platinum, etc.)

c Inert electrode (graphite, platinum, etc.) or unreactive metal (e.g. copper, silver, gold)

d Inert electrode (graphite, platinum, or glass electrode, which allows gas to be bubbled in)

6 Lithium is a very strong reducing agent. Fluorine is a very strong oxidising agent. If they were mixed together, there would be a violent exothermic reaction.

4.4 The electrochemical series

4.4 CHECK YOUR LEARNING

1 To determine the position of the other half-cells in the electrochemical series by functioning as a reference cell against which to measure the voltages of other cells.

2 The E^0 values change. Hence, the order of the half-equations in the electrochemical series can change, and the potential difference between cells can also change.

3 **a** 1.10 V

b 2.70 V

c 1.92 V

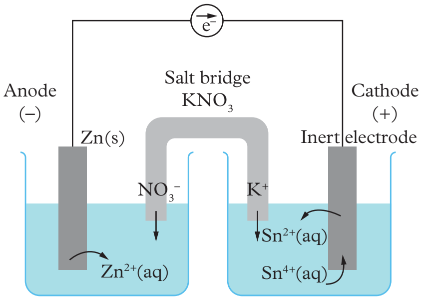
d 0.78 V

e 1.60 V

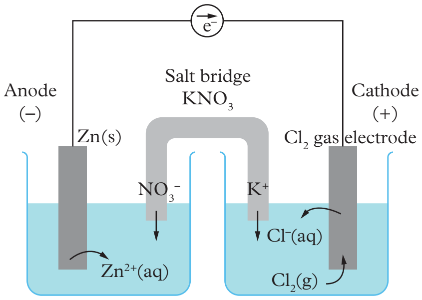
f 0.32 V

4 **a** $Sn^{4+}(aq)/Sn^{2+}(aq)$ and $Zn(s)/Zn^{2+}(aq)$: 0.91 V
$Zn(s)/Zn^{2+}(aq)$ and $Cl^-(aq)/Cl_2(g)$: 2.12 V
$Cl^-(aq)/Cl_2(g)$ and $Sn^{4+}(aq)/Sn^{2+}(aq)$: 1.21 V

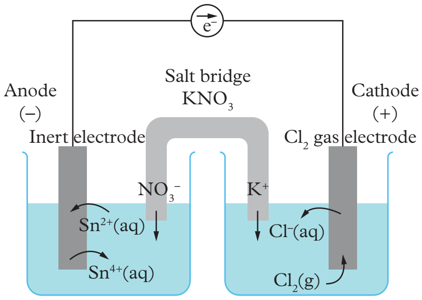
b $Sn^{4+}(aq)/Sn^{2+}(aq)$ and $Zn(s)/Zn^{2+}(aq)$
Reduction at the cathode:
$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$
Oxidation at the anode:
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
Overall redox reaction:
$Zn(s) + Sn^{4+}(aq) \rightarrow Zn^{2+}(aq) + Sn^{2+}(aq)$



$Zn(s)/Zn^{4+}(aq)$ and $Cl^-(aq)/Cl_2(g)$
Reduction at the cathode:
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$
Oxidation at the anode:
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
Overall redox reaction:
$Zn(s) + Cl_2(g) \rightarrow Zn^{2+}(aq) + 2Cl^-(aq)$



$Cl^-(aq)/Cl_2(g)$ and $Sn^{4+}(aq)/Sn^{2+}(aq)$
Reduction at the cathode:
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$
Oxidation at the anode:
$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq)\ 2e^-$
Overall redox reaction:
$Sn^{2+}(aq) + Cl_2(g) \rightarrow Sn^{2+}(aq) + 2Cl^-(aq)$



c The strongest oxidising agent is chlorine. The strongest reducing agent is zinc.

d The highest voltage is 2.12 V and is produced by the $Zn(s)/Zn^{2+}(aq)$ and $Cl^-(aq)/Cl_2(g)$ cell.

5 If a copper cell were used as the reference cell, all the voltages would change by +0.34 V.
The half-cells in the series would not change positions, as the copper cell can work as an oxidising agent or reducing agent, as hydrogen does when it is a reference cell, depending on what it is being tested against.

6 Magnesium and zinc are more reactive that iron, so they will oxidise in preference to the iron. They are stronger reducing agents than iron.

7 Student answers will vary but could suggest that the silver ions in the silver half-cell could react with the chloride ions from the salt bridge to form a silver chloride precipitate, breaking the internal circuit. The precipitation reaction can be prevented by replacing the sodium chloride salt bridge with a potassium nitrate (KNO_3) salt bridge. Unlike silver chloride, silver nitrate is soluble in water.

4.5 Features of fuel cells

4.5 REAL-WORLD CHEMISTRY

1 Reduction: $O_2(g)+2H_2O(l)+4e^- \rightarrow 4OH^-(aq)$
Oxidation: $H_2(g)+2OH^-(aq) \rightarrow 2H_2O(l)+2e^-$
2 $2H_2(g)+O_2(g) \rightarrow 2H_2O(l)$
3 In external tanks, oxygen gas and hydrogen gas are in direct contact with each other and react spontaneously, so they do not produce electrical energy but heat, light and sound energy.

4.5 CHECK YOUR LEARNING

1 Primary cells: contain a limited amount of reactants, only produce low voltages, and the zinc anode is oxidised and breaks down to form zinc ions in solution (e.g. a non-rechargeable alkaline battery).

Fuel cell: can produce an electrical current continuously if it has reactants supplied, contains inert electrodes (porous carbon) that do not break down, and have high-density power and are highly efficient, so they can power a car (e.g. a hydrogen–oxygen fuel cell).

2 The electrodes are constructed from porous carbon because this increases their surface area and therefore provides more opportunity for the gaseous reactant to come in contact with the electrodes.

3 Find the full labelled drawing in your obook pro.
  a Oxidation: $C_2H_5OH(l)+3H_2O(l) \rightarrow 12H^+(aq)+2CO_2(g)+12e^-$
  Reduction: $3O_2(g)+12H^+(aq)+12e^- \rightarrow 6H_2O(l)$
  b Reactants: C_2H_5OH, O_2
  Products: H_2O, CO_2
  Electrolyte: H_2SO_4
  c Oxidation occurs at the anode and reduction at the cathode.
  d Electrons: Anode to cathode
  Ions: Cations to cathode, anions to anode
  e $OC_2H_5OH(l)+3O_2(g) \rightarrow 3H_2O(l)+2CO_2(g)$

4 Find the full labelled drawing in your obook pro.
  a Oxidation: $CH_3OH(l)+6OH^-(aq) \rightarrow CO_2(g)+5H_2O(l)+6e^-$
  Reduction: $2H_2O(l)+O_2(g)+4e^- \rightarrow 4OH^-(aq)$
  b Reactants: CH_3OH, O_2
  Products: H_2O, CO_2
  Electrolyte: KOH
  c Oxidation occurs at the anode and reduction at the cathode.
  d Electrons: Anode to cathode
  Ions: Cations to cathode, anions to anode
  e Overall: $2CH_3OH(l)+3O_2(g) \rightarrow 2CO_2(g)+4H_2O(l)$

5 Find the comparison table in your obook pro.

4.6 Faraday's Laws

4.6 SKILL DRILL

1 IV: different half-cells
DV: the mass of each metal deposited at the cathode
2 If the same amount of electricity is passed through the electrolyte in four different half-cells, then the mass of the metal deposited at the cathode will be proportional to the amount of electrons required to reduce the metal ion.
3 To verify Faraday's second law by determining the mass of metal deposited at the cathode using four different half-cells compared to the standard hydrogen half-cell.

4.6 CHECK YOUR LEARNING

1 The total charge carried by 1 mol (6.02×10^{23}) of electrons; 96 500 C.
2 Electric charge (Q) is the quantity of electricity that passes through the cell. It is calculated using the formula $Q = I\times t$ where Q is the electric charge (in coulombs, C), I is the current passing through the cell (in amps, A) and t is the time for which the current flows (in seconds, s).
3 Faraday's second law ($Q = n(e^-)\times F$ (where $F = 96\,500$ C)) can be used to determine the number of moles that must be consumed to produce one mole of metal or ion. This means that 1, 2, 3 or another whole number of moles of electrons must be consumed to produce 1 mole of a metal.
4 4.6 min (2 sig fig)
5 1.2 A (2 sig fig)
6 **a** 0.33 L (2 sig fig)
  b 78% (2 sig fig)
7 **a** The ability to convert the reactants into products, and the products back into reactants. At 100% capacity, all of the reactants can be converted to products and back.
  b 2.2 A (2 sig fig)
  c 576.9 g or 6×10^2 g (1 sig fig)
8 **a** Reduction: $F_2(g)+2e^- \rightarrow 2F^-(aq)$
  Oxidation: $Li(s) \rightarrow Li^+(aq)+e^-$
  Overall: $F_2(g)+2Li(s) \rightarrow 2F^-(aq)+2Li^+(aq)$
  b 5.91 V; the battery is capable of producing the 5.5 V required.
  c 62 g (2 sig fig)
  d 111 L or 1.1×10^2 L (2 sig fig)
  e Student answers will vary.

4.7 Meeting society's energy needs

4.7 CHECK YOUR LEARNING

1 With the global population increasing, the demand for energy also increases, and this increases the demand for the finite reserves of fossil fuels. The burning of fossil fuels releases carbon dioxide and water into the atmosphere, which contributes to the greenhouse effect and climate change.
2 Fuel cells are energy-efficient due to their porous and catalysed electrodes, and because they can be designed to use fuels from sustainable sources. They convert chemical energy directly into electrical energy and are 60% efficient. A coal-fired power plant generates electrical energy from coal through several, less efficient energy transformations.
3 Use of renewable feedstocks: Raw materials or feedstocks should be made from renewable (mainly plant-based) materials, rather than from fossil fuels whenever practicable.
4 This ensures that resources are consumed and produced sustainably, which is Sustainable Development Goal 12: Responsible consumption and production. It seeks to prevent waste instead of treating or cleaning it up after it has been created, which is the green chemistry principle: Prevention. It also shifts away from a linear economy to a circular economy, because the focus is on using and reusing resources repeatedly, ensuring that they are available in the future.
5 Student answers will vary but should include an explanation of the generation, capture and safety of a fuel.
6 In fuel cells, the porous carbon electrodes may also be embedded with a catalyst which increases the rate of the reaction and therefore the efficiency of the fuel cell. This relates to the green chemistry principle: Catalysis: In green chemistry, the aim of this would be to generate the same product with less energy or waste in a reaction.

Chapter 4 review

MULTIPLE CHOICE

1	D	**2**	C	**3**	B	**4**	A	**5**	B
6	D	**7**	C	**8**	B	**9**	D	**10**	A

SHORT ANSWER

11 Oxidation state helps determine the number of electrons that have been gained or lost by an atom. The charge of an ion is determined by the number of electrons and protons.

12 Corrosion involves the reaction of a metal with oxygen. The term *oxidation* was originally used to describe reactions in which an element combines with oxygen.

13 Galvanic cells use a spontaneous chemical reaction to generate electricity when the half-cells are connected by a wire and salt bridge. If you combined the reactants of the cells together in one beaker, the spontaneous reaction would create heat energy instead of electrical energy.

14 This means the E^0 values in the electrochemical series apply and the voltage can be predicted, as well as the strongest reducing agent and oxidising agent.

15 An electrolyte is an ionic solution or gel that is capable of carrying charge. In a galvanic cell, the electrolyte often contains reactants for the spontaneous redox reaction. A salt bridge allows the flow of cations to the cathode and anions to the anode and completes the galvanic cell circuit. Salt bridges are only used to complete the circuit of galvanic cells set up in the laboratory for experimental purposes. In primary, secondary and fuel cells, electrolytes are used instead of salt bridges because salt bridges would dry out and be impractical.

16 The position of the elements in the periodic table can be used to identify trends such as metallic character and electronegativity. *Metallic character* describes how readily an atom loses an electron, whereas *non-metallic character* describes how readily an atom gains an electron. Metallic character increases as you move down a group and decreases as you move from left to right across a period. Non-metallic character increases as you move from bottom left of the table diagonally up to top right. Electronegativity is the tendency for an atom to attract electrons when in a chemical bond. Electronegativity tends to decrease as you move down a group and increase as you move from left to right across a period.

17 The oxidising agent is responsible for causing oxidation but itself undergoes reduction. The reducing agent is responsible for causing reduction but itself undergoes oxidation. Redox reactions involve a transfer of electrons from the reducing agent to the oxidising agent.

18 Reduction: $Pb^{2+}(aq)+2e^-\rightarrow Pb(s)$
Oxidation: $Fe(s)\rightarrow Fe^{2+}(aq)+2e^-$
Overall: $Fe(s)+Pb^{2+}(aq)\rightarrow Fe^{2+}(aq)+Pb(s)$

19 Oxidation: $Ni(s)\rightarrow Ni^{2+}(aq)+2e^-$
Reduction: $Cu^{2+}(aq)+2e^-\rightarrow Cu(s)$
Overall: $Ni(s)+Cu^{2+}(aq)\rightarrow Ni^{2+}(aq)+Cu(s)$

20 **a** –3 **b** +4
c +6 **d** +4

21 **a** +3 **b** +2
c +2 **d** 0

22 **a** $C:+4\rightarrow +2$
$H:0 \rightarrow +1$
$O:-2 \rightarrow -2$

b Hydrogen has been oxidised; carbon has been reduced.

c Hydrogen is the reducing agent; carbon dioxide is the oxidising agent.

d Oxidation: $H_2(g)\rightarrow 2H^+(aq)+ 2e^-$
Reduction: $CO_2(g)+2H^+(aq)+2e^-\rightarrow CO(g)+H_2O(g)$

e $CO_2(g)+H_2(g)\rightarrow CO(g)+H_2O(g)$

23 **a** $Zn:0 \rightarrow 2+$
$H:+1 \rightarrow 0$
$S:+6 \rightarrow +6$
$O:-2 \rightarrow -2$

b Zinc has been oxidised; hydrogen has been reduced.

c Sulfuric acid is the oxidising agent; zinc is the reducing agent.

d Oxidation: $Zn(s)\rightarrow Zn^{2+}(aq)+2e^-$
Reduction: $2H^+(aq)+2e^-\rightarrow H_2(g)$

e $Zn(s)+2H^+(aq)\rightarrow Zn^{2+}(aq)+H_2(g)$

24 **a** $Ag:0 \rightarrow +1$
$H:+1 \rightarrow 0$
$S:-2 \rightarrow -2$

b Silver has been oxidised; hydrogen has been reduced.

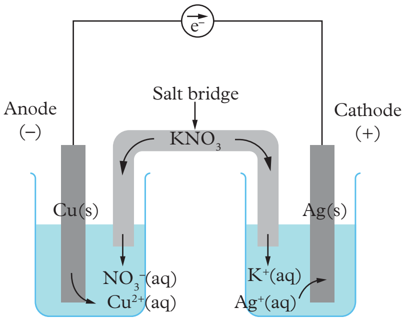
c Hydrogen sulfide is the oxidising agent; silver is the reducing agent.

d Oxidation: $Ag(s)\rightarrow Ag^+(aq)+e^-$
Reduction: $2H^+(g)+2e^-\rightarrow H_2(g)$

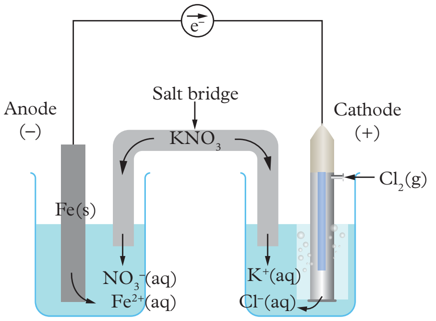
e $2Ag(s)+2H^+(g)\rightarrow 2Ag^+(s)+H_2(g)$

25 **a** Reduction: $Ag^+(aq)+e^-\rightarrow Ag(s)$
Oxidation: $Cu(s)^{2+}\rightarrow Cu^{2+}(aq)+2e^-$
Overall: $2Ag^+(aq)+Cu(s)\rightarrow 2Ag(s)+Cu^{2+}(aq)$

$E^0 = 0.46\ V$

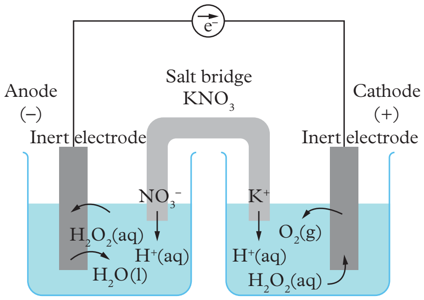


b Oxidation: $Fe(s)\rightarrow Fe^{2+}(aq)+ 2e^-$
Reduction: $Cl_2(g)+2e^-\rightarrow 2Cl^-(aq)$
Overall:
$Fe(s)+Cl_2(g)\rightarrow Fe^{2+}(aq)+2Cl^-(aq)$
$E^0 = +1.80\ V$



c Oxidation: $H_2O(aq)+2H^+(aq)+2e^-\rightarrow 2H_2O(1)$
Reduction: $H_2O_2(aq)\rightarrow O_2(g)+2H^+(aq)+2e^-$
Overall: $2H_2O(aq)\rightarrow O_2(g)+2H_2O(1)$

$E^0 = 1.09\ V$



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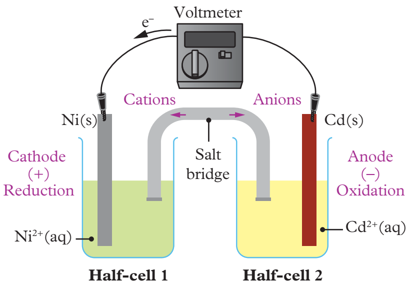


Oxidation: $Cd(s)\rightarrow Cd^{2+}(aq)+2e^-$
Reduction: $Ni^{2+}(aq)+2e^-\rightarrow Ni(s)$
Overall: $Cd(s)+Ni^{2+}(aq)\rightarrow Cd^{2+}(aq)+Ni(s)$
Salt bridge solution: KNO_3, NaCl or $NaNO_3$
$E^0 = 0.15\ V$

27 **a** **i** $Mg(s)$
ii $Fe^{2+}(aq)$
iii Oxidation:
$Mg(s)\rightarrow Mg^{2+}(aq)+2e^-$
Reduction:
$Fe^{2+}(aq)+2e^-\rightarrow Fe(s)$
iv $Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$

b **i** $Cu(s)$

ii $Ag^+(aq)$

iii Oxidation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

iv $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

c **i** $Al(s)$

ii $H^+(aq)$

iii Oxidation:

$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

Reduction:

$2H^+(aq) + 2e^- \rightarrow H^2(g)$

iv $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2\ (g)$

d **i** $H_2(g)$

ii $Fe^{3+}(aq)$

iii Oxidation:

$2OH^-(aq) + H_2(g) \rightarrow 2H_2O(l) + 2e^-$

Reduction: $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$

iv $2OH^-(aq) + H_2(g) + 2Fe^{3+}(aq) \rightarrow 2H_2\ O(l) + 2Fe^{2+}(aq)$

28 **a** i Yes, silver will be coated on the iron nail.

ii No, magnesium will not be coated on the iron nail.

iii Yes, copper will be coated on the iron nail.

iv Yes, lead will be coated on the iron nail.

v No, zinc will not be coated on the iron nail.

b The solutions containing Ag^+, Cu^{2+}, and Pb^{2+} will coat the nail because they are above iron in the electrochemical series and the ions will be reduced into their solid metal forms. In these cases, $Fe(s)$ is oxidised to Fe^{2+}. The solutions that will not coat the nail are Mg^{2+} and Zn^{2+} because they are below iron on the electrochemical series, and they will not react spontaneously with Fe.

c These are spectator ions in the reactions.

29 Student answers will vary, but an example summary can be found in your obook pro.

30 **a** Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

b Zinc electrode will lose 0.076 g.

c Copper electrode will gain 0.15 g.

31 **a** Reduction: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

Oxidation: $H_2\ Se(g) \rightarrow Se(s) + 2H^+(aq) + 2e^-$

b 4.6 L (2 sig fig)

32 5.8 L (2 sig fig)

33 **a** Oxidation: $Na(s) \rightarrow Na^+(aq) + e^-$

Reduction: $2H_2O(l) + 2e^- \rightarrow 2OH^-(aq)$

b $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq)$

c A strength is that sodium hydroxide can be produced by the spontaneous redox reaction between sodium metal and water; however, it is the not the best method to produce sodium hydroxide. Sodium metal can ignite spontaneously in air or water.

34 The subshell electron configuration of yttrium is $4d^15s^2$, so it will be found as the Y^{3+} ion in the compound yttrium barium copper oxide. Yttrium has an oxidation number of +3. Barium will have an oxidation number of +2. Oxygen will have an oxidation number of –2.

Two Cu atoms must have an oxidation number of +2 and one Cu atom must have an oxidation number of +3. In normal circumstances, it would be expected that Cu would have either a +1 or +2 oxidation number.

35 The available reactants are: $Ag(s)$, $Cu^{2+}(aq)$ and $SO_4^{2-}(aq)$. The sulfate ions are spectator ions so they do not react.

$Ag(s)$ is not a strong enough reducing agent to cause the reduction of $Cu^{2+}(aq)$ to $Cu(s)$. Similarly, $Cu^{2+}(aq)$ is not a strong enough oxidising agent to cause the oxidation of $Ag(s)$ to $Ag^+(aq)$. Therefore, no reaction will occur and hence no colour change will be observed.

36 Zinc is a more reactive metal than iron so will act as a sacrificial anode and undergo oxidation in preference to iron. Therefore, it can protect the steel from undergoing corrosion and forming rust.

37 **a** Both the Ag^+ and Au^+ solutions have a spontaneous reaction with the unknown metal and a layer of silver or gold forms on the unknown metal. Thus, Ag^+ and Au^+ must be above the unknown metal in the electrochemical series.

The unknown metal must be above Mg^{2+} and Al^{3+} on the electrochemical series because Mg and Al do not coat the unknown metal. As they are not reduced, Mg^{2+} and Al^{3+} must be weaker oxidising agents relative to *X*.

b The metal could be any on the electrochemical series between silver and aluminium All of these will be oxidised (hence the reducing agent) with silver and gold ions, but not react spontaneously with the magnesium or aluminium ions.

Unit 3 Area of Study 1 Checkpoint

MULTIPLE CHOICE

1 C **2** D **3** C **4** B **5** B

6 C **7** D **8** B **9** D **10** B

SHORT ANSWER

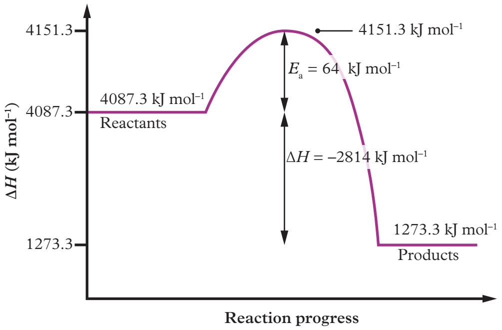
1 **a** Petrol

b A smaller amount of fossil fuel is used. This is a shift towards the use of renewable sources of fuel.

c Biodiesel is made from plant material that can be regrown within a lifetime. It is carbon-neutral. Petrodiesel is non-renewable, because it is made from fossil fuels, which cannot be regenerated within a lifetime. The CO_2 released is not reabsorbed and therefore remains in the atmosphere.

2 **a** –2814.0 kJ mol⁻¹ (5 sig fig)

b



c $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$

$\Delta H = -72\ kJ\ mol^{-1}$

d 57 674 mL or 57.7 L (3 sig fig)

3 **a** 39.0 J °C⁻¹

b –29.2 kJ mol⁻¹ (3 sig fig)

c The experimental value is lower that the theoretical value.

d Student answers will vary, but one example is that heat may have been lost to the surrounding environment through the lid of the calorimeter, or because of poor insulation.

4 **a** $Ni(s)$

b $Ni^{2+}(aq)/Ni(s)$

c The permanganate (MnO_4^-) and sulfate (SO_4^{2-}) half-cells both undergo reduction when combined with the hydrogen half-cell, and there is no data that verifies which would be the strongest oxidising agent.

d Student answers will vary, but may suggest attaching the permanganate (MnO_4^-) and sulfate (SO_4^{2-}) half-cells. Electrons flow from anode to cathode, so the strongest oxidising agent will be reduced at the cathode.

5 **a** positive / +

b →

c $CH_4(g) + 2H_2O(l) \rightarrow CO_2(g) + 4H^+(aq) + 4e^-$

d Platinum acts as a catalyst, increasing the rate of the reaction/decreasing the activation energy to keep temperatures low.

- **e** Fuel cells use catalysts on electrodes to increase the rate of the reaction, and have porous electrodes to increase their surface area and therefore the rate of the reaction.

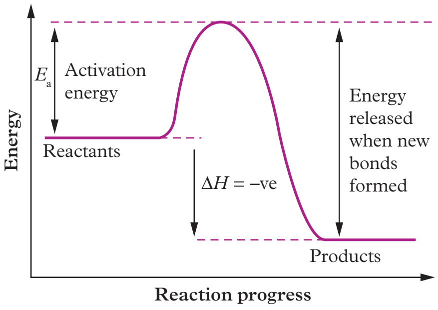
6 **a** 744 kJ/100 g (3 sig fig)
- **b** 604 kJ/100 g (3 sig fig)
- **c** Both theoretical values were lower than the theoretical values.
- **d** The results are not accurate because they are not close to the theoretical values. Precision cannot be commented upon as the experiment wasn't repeated.

Chapter 5: Rates of chemical reactions

GROUNDWORK

5A Activation energy is the minimum amount of energy required for a reaction to occur successfully.

5B



5.1 Factors affecting chemical reactions

5.1 REAL-WORLD CHEMISTRY

1 Powdered sugar has a larger surface area, resulting in more sugar 'available' to react. Therefore, this increases the frequency and proportion of successful collisions in the reaction.

2 Future plants should store sugar in cubes to reduce the surface area of the reactant. Sugar should be stored in sealed containers so that it does not react with the oxygen in the surroundings.

3 The dust should have been disposed of immediately after it was collected. Another safety precaution would be making just enough sugar to sell, rather than storing the sugar in bulk, which can build up more sugar dust in the air.

5.1 CHECK YOUR LEARNING

1 Collision theory states that reactants must collide in order to react. For a successful collision, reactants must collide in the correct orientation, or in the correct position, and with a minimum amount of energy to rearrange and form products.

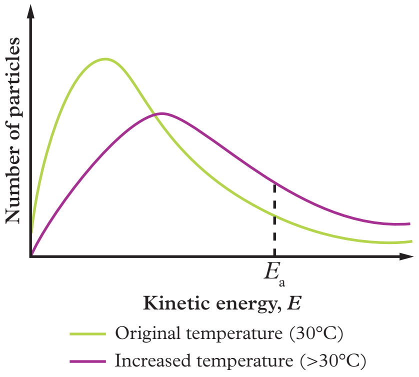
2 Decreased temperature, agitation, pressure, concentration, surface area and no catalysts

3 Activation energy is the minimum amount of energy required in a collision for a reaction to occur. If the energy contained within the bonds of the reactants is relatively small (i.e. the bonds are weak), a lower activation energy is required. Strong bonds contain larger amounts of energy, are harder to break apart and require a higher activation energy.

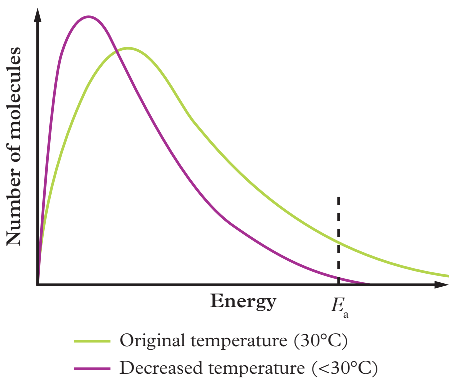
4 Only an increase in temperature will increase the frequency of collisions and proportion of successful collisions, as more particles will have energy greater than the activation energy.

5 **a** Increasing the concentration of the reactions will increase the number of collisions that can potentially occur. Simultaneously, it increases the pressure of the system as there is an increase in the number of particles, thereby increasing the rate of reaction. By increasing the temperature, the kinetic energy of each particle increases, speeding up the reaction.

b **i**



ii



- **c** The peak of the curves refers to the kinetic energy that is experienced by the highest number of molecules. It also represents the average kinetic energy of the reaction and is equivalent to the temperature of the system.
- **d** Increasing the temperature increases the kinetic energy experienced by the highest number of molecules in the system. It also increases the number of molecules that have energy after the activation energy has been reached.

6 The peak of each curve seems to appear at the same energy level, which implies that the temperature has not changed. The peak height for sample 2 is much shorter than the peak height for sample 1 – i.e. there is a smaller number of molecules. This could have happened due to a decrease in the concentration of the sample being used in the reaction mixture. Sample 1 is likely to have a faster rate of reaction since there would be more molecules with energy greater than the activation energy.

7 The manufacturer should have the car parts wrapped in suitable plastic to shield them from the rain, or take the car parts out of the rain and place them inside the building for protection.

8 Reactions occur at the surface of a solid reactant. Therefore, the greater the surface area, the greater the number of collisions. Smaller solids have larger surface areas. Therefore, the rate of dissolving the granulated sugar in a hot tea or coffee is higher than that for sugar cubes because of the increased surface area of the granulated sugar compared to that of the sugar cubes.

5.2 The role of catalysts

5.2 SKILL DRILL

1 IV concentration of H_2O_2
DV: mass of gas lost every 10 seconds

2 Chemicals used, amount of water, room temperature (18.0°C), balance used, measuring equipment (e.g. bulb pipette)

3 Not all control variables can be controlled, especially temperature. An exothermic reaction means that the reaction is releasing heat energy into the environment. This will alter the temperature of the room where the experiment is taking place.

4 Reliability: reliability is low because the repeated results varied significantly.

Validity: the experiment is valid because it still allows students to investigate the rate of decomposition of hydrogen peroxide.

Precision: because the results vary significantly, this experiment has low precision.

5.2 CHALLENGE

1 Homogeneous catalyst: aqueous state

Heterogeneous catalyst: solid state

2 Homogeneous catalyst

3 The homogeneous catalyst mixes the catalyst with the reactants, which allows for the reactants to easily bond with the surface of the catalyst. This allows for a faster rate of weakening or breaking the bonds contained within the reactants.

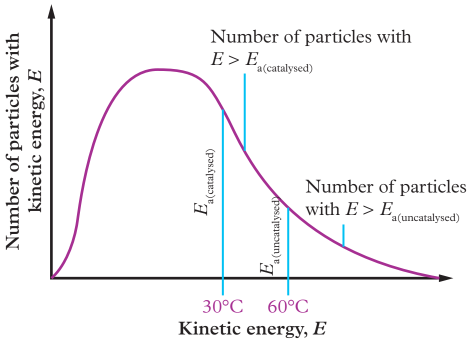
5.2 CHECK YOUR LEARNING

1 A catalyst is a substance that lowers the activation energy of a reaction and increases the rate of reaction without participating in the chemical reaction.

2 A homogeneous catalyst is a catalyst that is in the same state as the reactants, and a heterogeneous catalyst is a catalyst that is in a different state from the reactants.

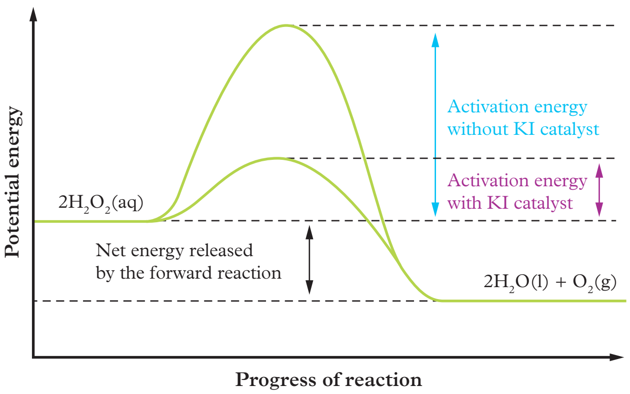
3 The catalyst only affects the activation energy of the reaction and does not affect the enthalpy of the reaction.

4



5 The fully labelled energy profile diagram below shows the dissociation of hydrogen peroxide (H_2O_2) using potassium iodide (KI) as a catalyst.

$$2H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(l)$$



6 a The production of ethanol by the fermentation of sugar:

$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$

The production of ethanol by the reaction of water with ethene:

$CH_2{=}CH_2 + H_2O \rightarrow CH_3CH_2OH$

b Reaction 1: enzymes of the yeast (e.g. zymase) and bacteria; reaction 2: phosphoric acid (H_3PO_4)

c The fermentation of sugar: Advantages: uses a renewable source (plant material), uses less energy and lower temperatures; Disadvantages: slow-batch process, produces CO_2 as waste; is naturally highly combustible if not stored properly

The hydration of ethene: Advantages: no waste products, produces ethanol continuously, is relatively cheap and economical, does not take as much time to produce. Disadvantages: uses a non-renewable source, requires a lot of energy for high-heat and high-pressure conditions.

7 a $Fe(NO_3)_3$ will cause a faster reaction because it can be mixed through the peroxide, causing the bonds to weaken or break at a quicker rate.

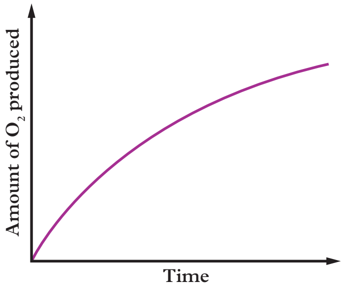
b The amount of catalyst used, the amount of peroxide, room temperature

c The environment should be a more controlled environment.

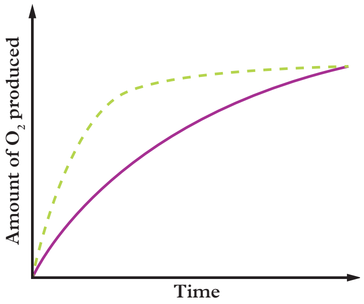
d The precision can be increased by keeping the controlled variables the same, so that the experiment can be conducted in the same conditions.

e 4.8 g (2 sig fig)

f



g



Chapter 5 review

MULTIPLE CHOICE

1	C	**2**	A	**3**	D	**4**	B	**5**	C
6	A	**7**	A	**8**	B	**9**	C	**10**	B

SHORT ANSWER

11 *Rate of a chemical reaction* refers to how fast the reaction takes place. This can be measured as a change (such as mass decrease) over time.

12 *y*-axis = concentration; *x*-axis = time

13 The activation energy of the catalysed pathway is lower than the activation energy of the uncatalysed pathway.

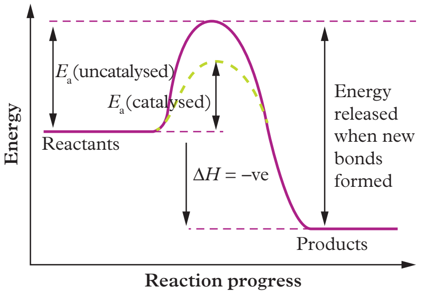
14 Increasing the temperature in an exothermic system, it will increase the rate of a chemical reaction because there is more energy provided for the reactants to form products.

15 Decreasing the surface area will decrease the rate of reaction, because it is harder for molecules to interact when less surface area is accessible to other reactants, resulting in less collisions.

16 Increasing the concentration of the reactants increases the number of particles in the system, so the rate of reaction is increased.

17 A catalyst is a substance that helps speed up the reaction but does not take part in the chemical reaction itself.

18



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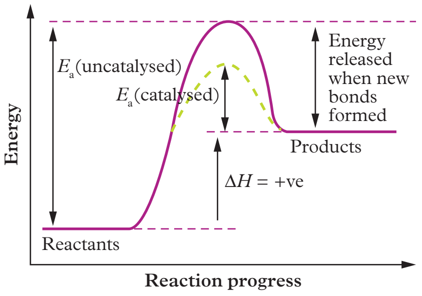
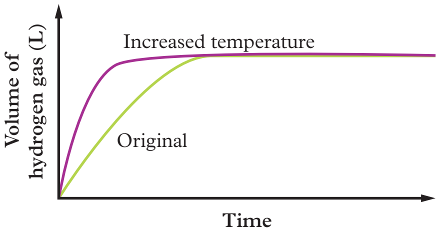


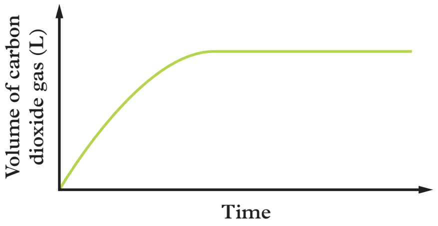
20 To increase the rate of a reaction, the activation energy must be decreased, or the number of successful collisions must be increased. Reactants must collide with a minimum amount of energy and with the correct orientation to rearrange and form products.

21

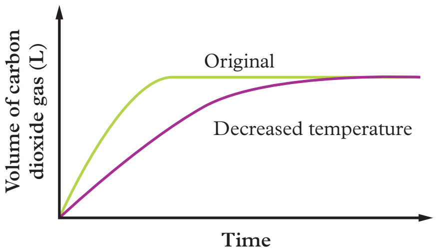
a



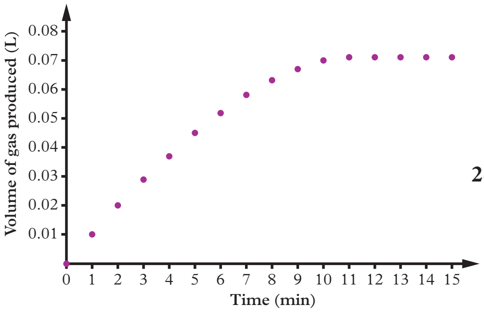
b



c



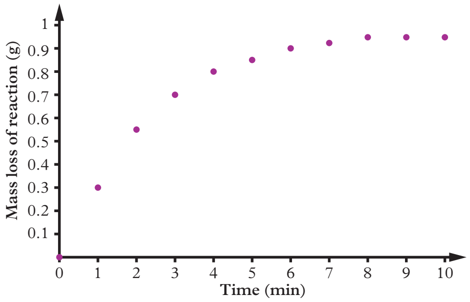
22 The rate of reaction for combustion could be measured by collecting the volume of CO_2 gas. The CO_2 gas generated could be measured by using the displacement of a known volume of water. The water is displaced as the gas enters the container of water, enabling the volume of gas generated to be measured. A graph of volume (on the *y*-axis) against time (on the *x*-axis) is then plotted (see the graph below).



This is gathering data by quantitative means.

23 The rate of reaction for precipitation could be measured by measuring:

- the mass of reactant loss over time. Quantitatively, the rate of a chemical precipitation reaction can be measured as a graph of the change in a reaction over time. The graph below represents the mass of a reactant lost over time. The graph demonstrates that the rate of the chemical reaction starts faster, and gradually slows over time.

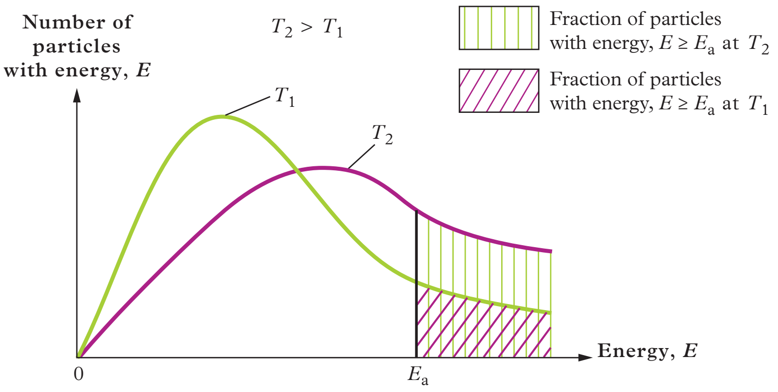


- qualitatively, the formation of precipitate formed over time. A timer could be used to time the precipitation reaction – that is, the time it takes for the cloudy precipitate to form and cover a black cross drawn on paper that is placed underneath the flask.

24 **a** The gas can be collected via a gas syringe that is attached to the opening of the system by a hose.

b A stopwatch is used to measure the amount of time needed to collect the amount of gas in the syringe.

25



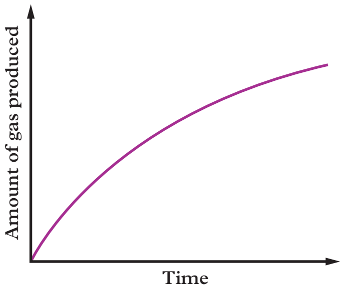
where T_1 = 80°C and T_2 = 120°C

26 a $Zn(g)+2HCl(aq) \rightarrow ZnCl_2(aq)+H_2(g)$

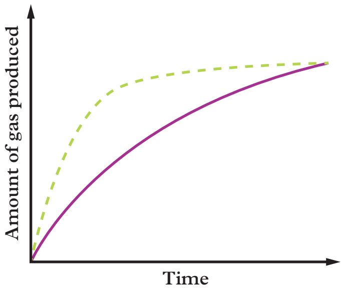
b 0.020 g (2 sig fig)

c 0.50 L (2 sig fig)

d



e

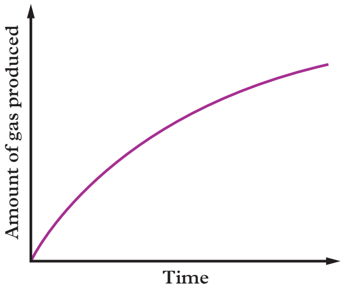


27 a $CaCO_3\ (s)+2HCl(aq) \rightarrow CaCl_2\ (aq)+ CO_2\ (g)+ H_2\ O(l)$

b 0.24 g (2 sig fig)

c 0.14 L (2 sig fig)

d



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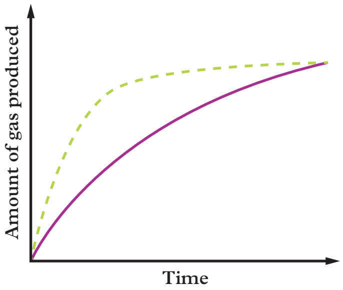
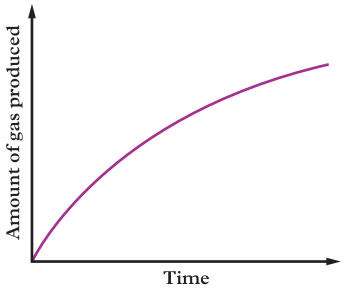


28 a 2.8 g (2 sig fig)

b 2.0 L (2 sig fig)

c



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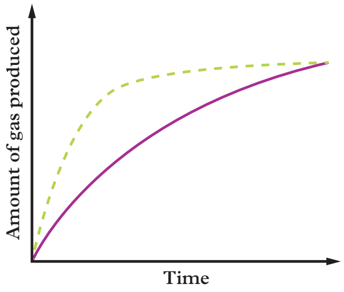
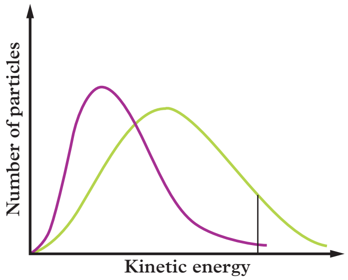


29 a Kinetic energy experienced by the highest number of particles in the reaction

b More particles in the system have increased kinetic energy, giving a broader curve. It is lower than the red curve because the distribution of kinetic energy experienced by the highest number of particles in the system has changed.

c The purple curve

d



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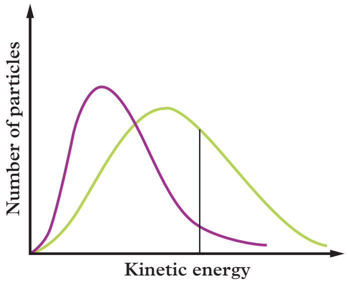


f Increasing the temperature increases the kinetic energy of the particles in the reactants, thus increasing the rate of reaction.

g Adding a catalyst speeds up the reaction by reducing the amount of energy required for successful collisions to occur (reduces the activation energy).

30 a By measuring the mass lost from the reactant (Mg(s)) over time and plotting a graph of mass against time

b Increased surface area of the cube and increased HCl concentration

c There are too many variables being trialled in the same reaction (reaction 2).

d Increased surface area (cube of Mg), and increased concentration of 2 M HCl, should each give a higher rate of reaction than with Mg ribbon and 1 M HCl, respectively.

e If there is excess HCl present, the magnesium will react to give a little more H_2 gas, but it will eventually stop reacting when the HCl is used; that is, the HCl becomes the limiting reactant. Therefore, adding more magnesium will not restart the reaction.

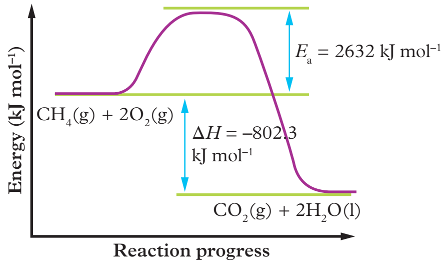
31 a Endothermic; the products have more energy than the reactants, and energy is absorbed from the environment as heat (quantity III).

b Quantity I

c When the reaction is switched from endothermic to exothermic, I < II, so III will be affected (ΔH is negative, exothermic). When the reaction is switched from exothermic to endothermic, I > II so III will be affected (ΔH is positive, endothermic).

d

Exothermic combustion of methane



32 a IV: temperature; DV: rate of reaction

b The higher the temperature, the faster the rate of reaction because the reactants would be provided with more kinetic energy, allowing for more frequent, successful collisions.

c Number of pieces/amount of magnesium, room temperature, measuring equipment

d Should be using exactly the same amount of magnesium; keep the time when recording the loss of gas the same; measure the time passed as the reaction starts and comes to a stop.

e Cannot be commented on. When the temperature increases, it is expected that the rate of reaction will increase. However, since the rate of reaction cannot be calculated, the student would not be able to compare it with the theoretical value.

f Low precision, because each trial yielded a different result

g Low validity, because the student would not be able to calculate the rate of experiment without measuring the time

Chapter 6: Extent of chemical reactions

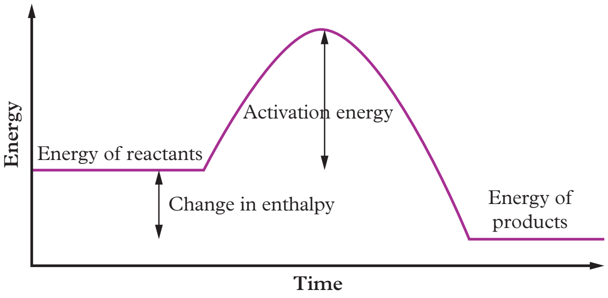
GROUNDWORK

6A Student answers will vary, but two examples include:

- increasing surface area of reactants: more particles are exposed for collisions and reactions to occur.
- increasing temperature: increases the kinetic energy and therefore the rate of collisions, as well as the likelihood of the collisions being above the activation energy.

6B As a chemical reaction occurs, the concentration of reactants decreases and the concentration of products increases.

6C



6.1 The rate and extent of reactions

6.1 SKILL DRILL

1 The reaction produces carbon dioxide gas, which has been lost to the environment and cannot be weighed.

2 0.7 g (1 sig fig)

3 The last repeat result is likely to be more accurate based on the calculated theoretical mass prior to rounding to significant figures.

4 The precision of the results is reasonably high because there is only a very small variation between repeats (±0.02 g). The repeatability and reproducibility of the data can be high if accurate measuring tools are used – this is not specified in the experiment outlined.

6.1 CHECK YOUR LEARNING

1 For a reverse reaction to occur, the original product particles must collide with one another with enough activation energy.

2 The rate of reaction is a measure of how fast a reaction occurs – the change in concentration of reaction or product per unit of time. The extent of reaction is a comparison of forward and backward reactions – a measure of the magnitude to which a reaction proceeds in the forward direction.

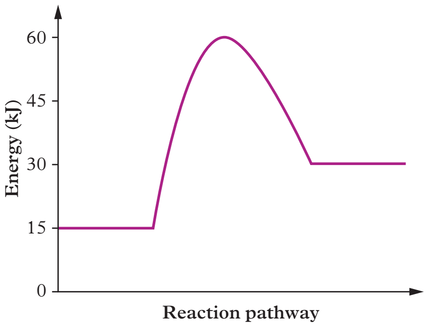
3 Reversible reactions in a closed system will always have all particles of reactants and products present because forward and backward reactions will continuously occur. Irreversible reactions in a closed system will have at least one of the reactants used up until there is nothing left (limiting agent). The reaction will only occur in the forward direction until at least one of the reactants is used up.

4 The rate of reaction will not have an impact on the yield.

5 **a** $-15\ \mathrm{kJ\ mol^{-1}}$

b $30\ \mathrm{kJ\ mol^{-1}}$

c



d $45\ \mathrm{kJ\ mol^{-1}}$

e The reaction is not reversible because the energy of the products formed from the forward reaction (40 kJ) is lower than the activation energy for the reverse reaction (45 kJ).

6.2 The nature of homogeneous equilibria

6.2 CHECK YOUR LEARNING

1 The reaction must be reversible. It must be in a closed system and the rates of the forward and reverse reactions become equal.

2 Homogeneous equilibrium systems have all reactants and products in the same state, which maximises the ability of the particles to collide successfully.

3 **a** Homogeneous
  b Heterogeneous
  c Heterogeneous
  d Homogeneous

4 $3H_2(g)+N_2(g) \rightleftharpoons 2NH_3(g)$

Before: the concentration of nitrogen and hydrogen will decrease as ammonia is formed by the above forward reaction (i.e. ammonia concentration will increase).

After: the concentration of all compounds will drop dramatically as the gases escape the smashed jar.

5 **a** **i** Reactants: $N_2O_5(g)$
  Products of forward reaction: $NO_2(g)$, $O_2(aq)$
  ii Reactants: $H_2(g)$, $I_2(g)$
  Products $HI(g)$
  iii Reactants: $SO_2(g)$, $O_2(g)$
  Products $SO_3(g)$
  iv Reactants: $NOCl(g)$
  Products $NO(g)$, $Cl_2(g)$
b **i** $2N_2O_5(g) \rightleftharpoons 4NO_2(g)+O_2(aq)$
  ii $H_2(g)+I_2(g) \rightleftharpoons 2HI(g)$
  iii $2SO_2(g)+O_2(g) \rightleftharpoons 2SO_3(g)$
  iv $NOCl(g) \rightleftharpoons NO(g)+Cl_2(g)$

6 The Earth is an open equilibrium system for matter. According to the law of conservation of mass, matter cannot be created or destroyed. Matter on Earth changes from one form to another, but the system is open because matter arrives on Earth from space as space dust and meteorites, and leaves as gases and space craft, for example.

6.3 Changes in the position of equilibrium

6.3 CHECK YOUR LEARNING

1 Le Chatelier's principle states that if an equilibrium system is subjected to a change, the system will adjust itself to partially oppose the effect of the change.

2 If the volume is increased or the pressure is decreased, the reaction will favour the direction that produces more particles. This will result in the increase of concentration of particles. If the

volume is decreased or the pressure is increased, then the reaction will favour the direction that produces fewer particles. This will result in the decrease of concentration of particles.

3 When the temperature is increased, the reaction will favour the direction that counteracts the temperature increase. If the temperature is increased in an endothermic reaction system, then the reaction will favour the forward reaction, resulting in a net forward reaction. If the temperature is increased in an exothermic reaction system, then the reaction will favour the backward reaction, resulting in a net reverse reaction.

4 **a** When volume is increased, pressure decreases – the equilibrium will move to the side with the largest number of particles. Therefore, it will shift towards the reactant side (there are three particles of reactants and only two on the product side) and there will be a net reverse reaction.

b When temperature is decreased for an exothermic reaction, there is a net forward reaction.

c Increase in product concentration means that the system will shift to partially oppose the addition, so there is a net reverse reaction.

d As this is an exothermic reaction, a temperature increase means there is a net reverse reaction.

5 Equilibrium reactions are dynamic where forward and backward reactions are usually occurring at the same rate. When a change is made to an equilibrium system that disrupts the balance, one of the reactions (forward or backward) will be favoured, resulting in a net reaction. This means that the other reaction is still occurring, but the net reaction is being favoured.

6 When there is a change in the equilibrium, the system will favour the direction of reaction that will partially oppose the external change, to minimise impact. The term *partially oppose* is used because the shift will only favour one reaction over the other, rather than cease the secondary reaction completely and the reaction will never reach the levels of the original equilibrium.

6.4 Using Le Chatelier's principle to increase yield

6.4 CHECK YOUR LEARNING

1 The yield is the quantity of the wanted product obtained from a chemical reaction.

2 When asked about the yield of a reaction, unless specified otherwise, the products of the net forward direction should be used.

3 According to Le Chatelier's principle, if the desired product was continuously removed, the equilibrium yield will be increased.

4 **a** Adding reactants (CO and O_2) into the vessel, removing products (CO_2) from the vessel, increasing the pressure of the vessel

b Adding reactants (N_2O_5) into the vessel, removing products (NO_2 and O_2) from the vessel, diluting the solution by adding water

c Adding reactants (H_3PO_4) into the vessel, removing products (PO_4^+ and H^+) from the vessel, diluting the solution by adding water

d Adding reactants (NO and H_2O) into the vessel, removing products (NH_3 and O_2) from the vessel, decreasing pressure of the vessel

5 **a** Adding reactants (SO_4 and O_2) into the vessel, removing products (SO_4) from the vessel, decreasing pressure of the vessel, decreasing the temperature of the vessel

b Adding reactants (N_2 and O_2) into the vessel, removing products (NO_2) from the vessel, decreasing pressure of the vessel, increasing the temperature of the vessel

c Adding reactants (NO_2 and O_2) into the vessel, removing products (N_2 and O_3) from the vessel, increasing the temperature of the vessel.

6 $2NO_2(g) \rightleftharpoons N_2O_4(g)$

By decreasing the volume initially, the concentrations of all products and reactants increases, so the colour instantly becomes a darker brown. However, the system will oppose this concentration change; because the pressure is increased, the system will shift to the side with the least particles. This means a shift towards the colourless N_2O_4, making the reaction mixture lighter in colour.

6.5 Equilibrium expressions

6.5 CHALLENGE

3.5 M^{-1} (2 sig fig)

6.5 CHECK YOUR LEARNING

1 If the temperature of an exothermic reaction increases, it will favour the reverse reaction, which will decrease the *K* value.

2 If the temperature of an endothermic reaction increases, it will favour the forward reaction, which will increase the *K* value.

3 **a** $K = \frac{[HI]^2}{[H_2][I_2]}$

b $K = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$

c $K = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$

d $K = \frac{[CO][H_3]^3}{[CH_4][H_2O]}$

4 **a** No units

b M

c M^{-1}

d M^{-2}

5 **a** Increase

b Decrease

c Decrease

6 $7.5 \times 10^3 M^{-2}$ (2 sig fig)

7 $7 \times 10^3 M^2$ (1 sig fig)

8 The equilibrium constants are not always the same due to the variations of the coefficients for each compound. This will change the powers of the concentrations when calculating the equilibrium constant.

9 The equilibrium constant is the ratio of reactants to products when the equation is at equilibrium. The more products formed, the larger the equilibrium constant.

6.6 The reaction quotient (*Q*)

6.6 SKILL DRILL

1 Student answers will vary but may include: changing the temperature of the reaction to determine whether the reaction is endothermic or exothermic or changing the concentration of the reaction to determines the rough particle ratio between products and reactants. The scientist can check whether the reaction has reached equilibrium by checking the concentration of substances at regular time intervals. The reaction has reached equilibrium if the concentrations no longer change.

2 Equilibrium reactions do not have a 100% yield. Changes to the process can be made to increase the reaction yield, such as regular addition of reactants into the vessel or regular removal of products from the vessel.

6.6 CHECK YOUR LEARNING

1 The equilibrium constant can only be calculated with concentrations of compounds after the reaction has reached equilibrium. The reaction quotient can be calculated with concentrations of compounds at any moment in time. The reaction quotient also gives an indicator of how close the reaction is to equilibrium when compared to the equilibrium constant.

2 To check how far along the reaction has proceeded and how close it is to reaching equilibrium.

3 If the reaction constant and equilibrium constant are the same, then the reaction is at equilibrium.

4 If $Q > K$, the position of the equilibrium will shift to cause a net reverse reaction to form more reactants.

5 If $Q < K$, the position of the equilibrium will shift to cause a net forward reaction to form more products.

6 **a** The concentration of products is very small.

b There are significant amounts of both products and reactants.

c The concentration of reactants is very small.

7 **a** $16\,M^{-1}$ (2 sig fig)

b $Q = 16\,M^{-1}$ and $K = 2.0\times10^{-2}\,M^{-1}$, so the equilibrium must shift backward (reverse reaction will be favoured) to reach equilibrium.

8 **a** $0.67\,M^{-1}$ (2 sig fig)

b $Q = 0.667\,M^{-1}$ and $K = 2\,M^{-1}$, so the equilibrium must shift forward (the forward reaction will be favoured) to reach equilibrium.

6.7 Green chemistry and equilibrium

6.7 CHECK YOUR LEARNING

1 The rate of reaction will increase with temperature increase. However, the yield of products will also be reduced in exothermic reactions because the reaction will shift towards the reverse reaction to counteract the increase in temperature. Increasing temperatures above room temperature will require additional energy input, which will decrease energy efficiency and goes against the green chemistry principles.

2 When increasing the pressure, the yield will decrease if the forward reaction provides more particles than the reverse reaction. On the other hand, the yield will increase if the forward reaction provides fewer particles than the reverse reaction. Increasing pressure above regular pressure will require additional energy input, which will decrease energy efficiency and goes against the green chemistry principles.

3 The yield can be increased by adding reactants or removing products continuously to promote the forward reaction. The addition of a catalyst to an equilibrium reaction will not change the yield but will increase the rate of the reactions by decreasing the activation energy required.

4 **a** Increasing pressure, adding a catalyst, adding reactants and removing products. It may not be worth changing the temperature because of the trade-off.

b Adding a catalyst, adding reactants and removing products. It may not be worth changing the pressure or temperature because of the trade-offs.

c Increasing pressure, adding a catalyst, adding reactants and removing products. It may not be worth changing the temperature because of the trade-off.

5 **a** Adding a catalyst, adding reactants and removing products. It may not be worth changing the pressure or temperature because of the trade-offs.

b Increasing the surface area of the solids, adding a catalyst, adding reactants and removing products. It may not be worth changing the temperature because of the trade-off.

c Increasing pressure, adding a catalyst, adding reactants and removing products. It may not be worth changing the temperature because of the trade-off.

6 The green chemistry principle involves the use of catalysts to lower the activation energy of reactions to ensure that excess energy is not wasted to increase the rate of the chemical reaction through increasing temperature and pressure. A catalyst can be continuously reused and may decrease the need for excess reactant, which will in turn reduce waste. The addition of catalyst will increase the reaction rate but will have no impact on the yield of the reaction.

7 This green chemistry principle promotes having chemical reactions occur in room temperature and pressure to ensure that excess energy is not wasted and to increase the rate of chemical reaction through increasing temperature and pressure. The rate of reaction can be increased through the addition of a catalyst, and the yield can be increased by the removal of wanted products to promote forward reaction.

8
- Increasing pressure: net forward reaction, resulting in more product, increases concentration, giving a faster rate of reaction.
- Increasing temperature: increases the rate of reaction, but results in a net backward reaction, meaning there is a decrease in the amount of product.

In industry, it is expensive to increase pressure and too high a temperature will cause a loss of product, so the trade-off is to use 400°C and 200 atm, and an iron catalyst to increase the rate of reaction by lowering the activation energy of the reaction.

Chapter 6 review

MULTIPLE CHOICE

1 D **2** C **3** C **4** C **5** D
6 C **7** B **8** A **9** B **10** A

SHORT ANSWER

11 **a** $K = \dfrac{[NO_2]^2}{[N_2O_4]}$

b $K = \dfrac{[CH_3OH]}{[CO][H_2]^2}$

12 Adding a catalyst to an equilibrium system enables it to reach equilibrium more quickly. It does not increase or decrease the concentration of any chemical within the system and therefore does not cause a 'change', as described by Le Chatelier, so it does not affect the position of the equilibrium. Therefore, it does not result in a net forward or backward reaction and no extra product is formed.

13 Adding an inert gas to an equilibrium will have no effect on the position of the equilibrium. This is because the inert gas does not cause any change in the concentration of reactants or products in the container. The same number of reactants and product molecules exist within the same volume of space, regardless of the increase in pressure caused by the inert gas.

14 **a** K is between 10^{-4} and 10^{4}, which means that there is a significant concentration of both reactants and products in the vessel. This accounts for the light brown colour, which is a mixture of the colourless N_2O_4 and the dark brown NO_2 gases.

b When the volume of the system doubles, it decreases the pressure/concentration of the particles inside the vessel. This shifts the reaction to favour the forward direction, which produces more particles to partially oppose the change. This will make the gas darker in colour.

c As the reaction is an endothermic reaction, the increase in temperature will shift the reaction to favour the forward reaction. This net forward shift increases the amount of dark brown NO_2 gas while the colourless N_2O_4 gas is consumed, which will make the gas darker in colour.

d When a colourless inert gas is injected into the system, this will have no effect on the system.

15 **a** Ag^+ ions react with SCN^- ions to form a white precipitate. This decreases the SCN^- concentration and causes a net reverse reaction.

b The F^- ions from NaF react with Fe^{3+}, which decreases the Fe^{3+} concentration and causes a net reverse reaction.

c Dilution causes all concentrations to decrease. The reaction shifts to the side with the largest number of particles, so there is a net reverse shift.

d NaOH decreases the amount of Fe^{3+} as it forms a precipitate of $Fe(OH)_3$, so the reaction shifts to the left.

e Pb^{2+} ions react with SCN^- ions to form a white precipitate. This decreases the SCN^- concentration and causes a net reverse reaction.

16 The reaction is endothermic because when the temperature increased, so did the equilibrium constant. This means that there are higher quantities of products formed.

17 **a** The concentration of reactants is greater than the concentration of products.

b **i** An increase in product concentration will drive the reaction backward to partially oppose the change.

ii Decreasing the temperature will cause a net backward reaction because the reaction is endothermic.

iii An increase in volume will cause a decrease in pressure, which will favour the side with the greater number of particles and cause a net backward reaction.

18 At the time that $Q>K$, the ratio of products is higher than when the reaction is at equilibrium. This means that the reverse reaction is favoured, Q until it matches K.

19 At the time that $Q<K$, the ratio of products is lower than when the reaction is at equilibrium. This means that the reaction will favour the forward reaction is favoured, increasing Q until it matches K.

20 The term *partially oppose* is used because the shift will only favour one reaction over the other, rather than ceasing the secondary reaction completely and it will never reach the original variables.

21 0.19 M (2 sig fig)

22 **a** Reaction 1: no unit; reaction 2: M^{-1}

b **i** 2.3×10^{24}

ii $5.4\times10^{48}\,M^2$

iii 7.5×10^{-9}

iv $8.0\times10^{4}\,M^{-1}$

v $3.5\times10^{-3}\,M$

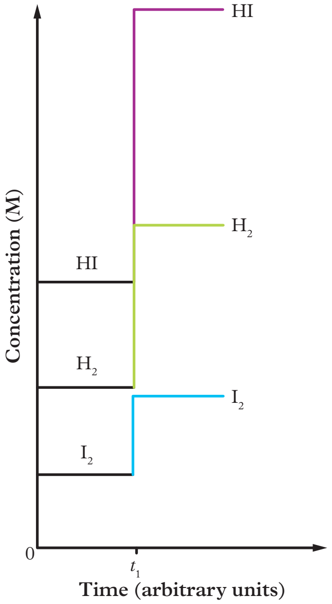
23 **a** 12.8

b Backward (reverse reaction)

c There are no units because the sum of the coefficients for reactants and products is the same.

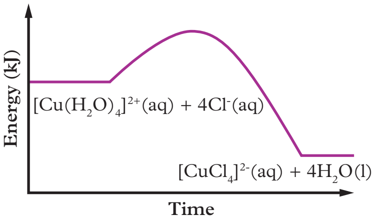
d When the volume of the system at equilibrium is halved, all the concentrations will be doubled. There will be no shift in the equilibrium system either forward or backward. This is because both the forward and backward reactions yield two particles each and will not be affected by changes in concentration/pressure.

e

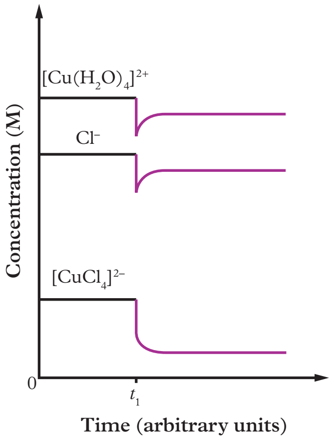


24 **a** The colour change indicates that there are higher amounts of products formed when the temperature is increased. This makes the reaction an endothermic reaction.

b



c



25 **a** The equilibrium constant for the equation is very large (greater than 10^4), which indicates that there is a much larger amount of the product (NO_2) compared to reactants. Therefore, it can be assumed that most of the 1.50 mol is still intact.

b 0.375 M (3 sig fig)

c 1.76×10^{-4} M (3 sig fig)

26 **a** The equilibrium constant for the equation is very small (smaller than 10^{-4}), which indicates that there is a much larger amount of the reactant (HI) compared to products. Therefore, it can be assumed that most of the 2 mol is still intact.

b 1.0 M (2 sig fig)

c 7.5×10^{-3} M (2 sig fig)

27 A buffer is a component of a solution that helps the solution to resist small pH changes. These buffers usually involve an equilibrium reaction. The pH in the human body needs to be tightly regulated and have various buffers in place to maintain the pH within suitable parameters, including the carbon dioxide dissolved in blood.

28 **a** $K = \dfrac{[NH_3]^2}{[N_2]\,[H_2]^3}$

b $0.47\,M^{-2}$ (2 sig fig)

c Forward

29 **a** The rate of ammonia formation was constant. This is likely to represent the reaction system at equilibrium.

b The rate of ammonia formation sharply increased. This could be due to a change in the system; for example, the addition of reactants, increase in temperature, increase in pressure or decrease in volume, or the addition of a catalyst.

c The rate of ammonia formation decreased. This is likely due to the system shifting to oppose the change in part **b**.

d The rate of ammonia formation became constant again. This is likely due to the system reaching a new equilibrium.

30 a Catalytic oxidation of ammonia:

$4NH_3(g)+5O_2(g) \rightleftharpoons 4NO(g)+6H_2O(g)$ $\Delta H = -907\,kJ\,mol^{-1}$ (exothermic)

Oxidation of nitrogen monoxide:

$2NO(g)+O_2(g) \rightleftharpoons 2NO_2(g)$ $\Delta H = -114\,kJ\,mol^{-1}$ (exothermic)

Absorption of nitrogen dioxide:

$3NO_2(g)+H_2O(l) \rightleftharpoons 2HNO_3(aq)+NO(g)$ $\Delta H = +126\,kJ\,mol^{-1}$ (endothermic)

b Reaction I: platinum/rhodium catalyst, high temperature (820–930°C), high pressure (1100 kPa)

Reaction II: low temperature, high pressure

Reaction III: the reaction vessels are packed with inert granular material to increase contact time between reactants; concentrations of 45–50% nitric acid are produced from these reactions.

c A trade-off, for example, is when a reaction requires low temperature for products to be formed for equilibrium reasons, but low temperatures slow the rate of reaction, so a compromise has to be reached. This reaction is unusual because the rate of reaction II increases as the temperature decreases. There is no conflict between the conditions needed for high equilibrium yield and economic reaction rate, because as it is cooled, the rate increases. Therefore, no real trade-off is required in the production of nitric acid by the Ostwald process.

Chapter 7: Production of chemicals using electrolysis

GROUNDWORK

7A Oxidation: $S^{2-}(aq) \rightarrow S(s)+2e^-$

Reduction: $Cl_2(g)+2e^- \rightarrow 2Cl^-(aq)$

7B $Cl^-(aq) > I^-(aq) > Fe(s) > Al(s)$

7C E^0 is the half-cell potential and is the potential difference of a cell under standard conditions.

7D Separate half-cells, salt bridge, galvanometer or light bulb

7.1 Electrolysis reactions

7.1 CHECK YOUR LEARNING

1 They require an external electrical energy source to provide energy for the reaction to occur.

2 To reduce competing reactions occurring in the cell at the same time

3 The electrochemical series is accurate only at standard conditions of 25°C, solutions with a concentration of 1.0 M and pressure of 1 atm. If non-standard conditions are used, E^0 may change. This can cause half-equations that have very similar E^0 to swap positions on the electrochemical series.

4 The electrons flow from the anode to the cathode in both types of cells. In a galvanic cell, the electrons flow from the negative to the positive electrode; in an electrolytic cell, they flow from the positive to the negative electrode.

5 **a** **i** Strongest oxidising agent: $K^+(l)$

Strongest reducing agent: $Br^-(l)$

ii Oxidation: $2Br^-(l) \rightarrow Br_2(l)+2e^-$

Reduction: $K^+(l)+e^- \rightarrow K(l)$

iii $2Br^-(l)+2K^+(l) \rightarrow Br_2(l)+2K(l)$

b **i** Strongest oxidising agent: $Li^+(l)$

Strongest reducing agent: $F^-(l)$

ii Oxidation: $2F^-(l) \rightarrow F_2(l)+2e^-$

Reduction: $Li^+(l)+e^- \rightarrow Li(l)$

iii $2F^-(l)+2Li^+(l) \rightarrow F_2(l)+2Li(l)$

c **i** Strongest oxidising agent: $H_2O(l)$

Strongest reducing agent: $H_2O(l)$

ii Oxidation: $H_2O(l) \rightarrow O_2(g)+4H^+(aq)+4e^-$

Reduction: $H_2O(l)+2e^- \rightarrow H_2(g)+2OH^-(aq)$

iii $2H_2O(l) \rightarrow 2H_2(g)+O_2(g)$

d **i** Strongest oxidising agent: $H_2O_2(aq)$

Strongest reducing agent: $H_2O_2(aq)$

ii Oxidation: $H_2O_2(aq) \rightarrow O_2(g)+2H^+(aq)+2e^-$

Reduction: $H_2O_2(aq)+2H^+(aq)+2e^- \rightarrow 2H_2O(l)$

iii $2H_2O_2(aq) \rightarrow O_2(g)+2H_2O(l)$

7.2 Features of commercial electrolytic cells

7.2 CHECK YOUR LEARNING

1 Galvanic cells have individual half-cells; electrolytic cells have a single container. Galvanic cells have a salt bridge; electrolytic cells do not. Galvanic cells use a galvanometer/lightbulb, while electrolytic cells require a power supply.

2 Reactive electrodes are preferred in the process of electroplating, where one metal is coated over another.

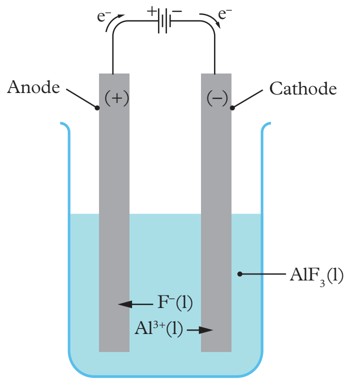
3 H_2O is a stronger reducing agent than O^{2-} and will react preferentially. Fe^{2+} is a stronger oxidising agent and will react preferentially. This will result in the products of $O_2(g)$ and $Fe(s)$ instead of the desired $CO_2(g)$ and $Al(s)$.

4 Conducting the Hall–Héroult cell in an aqueous solution will result in the addition of a competing reactant H_2O. Water is a stronger oxidising and reducing agent compared to the other reactants present in the cell. This will result in the formation of H_2 and O_2 instead of the desired $CO_2(g)$ and $Al(s)$.

Oxidation: $2H_2O(l) \rightarrow O_2(g)+4H^+(aq)+4e^-$

Reduction: $2H_2O(l)+2e^- \rightarrow 2H_2(g)+2OH^-(aq)$

5 **a**



Inert electrodes are made of platinum or graphite.

Strongest oxidising agent: $Al^{3+}(l)$

Strongest reducing agent: $F^-(l)$

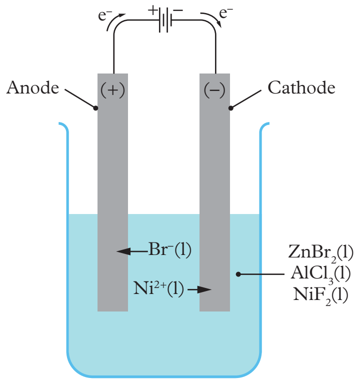
Oxidation: $2F^-(l) \rightarrow F_2(g)+2e^-$

Reduction: $Al^{3+}(l)+3e^- \rightarrow Al(s)$

Overall: $6F^-(l)+2Al^{3+}(l) \rightarrow 3F_2(g)+2Al(s)$

At least 4.53 V is required.

b



Inert electrodes are made of platinum or graphite.

Strongest oxidising agent: $Ni^{2+}(l)$

Strongest reducing agent: $Br^-(l)$

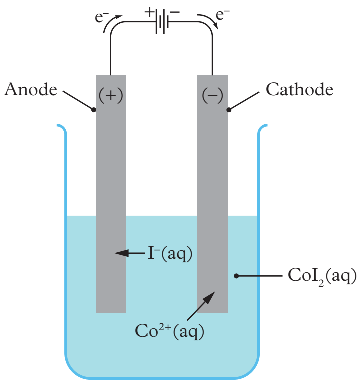
Reduction: $Ni^{2+}(l)+2e^- \rightarrow Ni(s)$

Oxidation: $2Br^-(l) \rightarrow Br_2(g)+2e^-$

Overall: $Ni^{2+}(l)+2Br^-(l) \rightarrow Br_2(g)+Ni(s)$

At least 1.34 V is required.

c



Inert electrodes are made of platinum or graphite.

Strongest oxidising agent: $Co^{2+}(aq)$

Strongest reducing agent: $I^-(aq)$

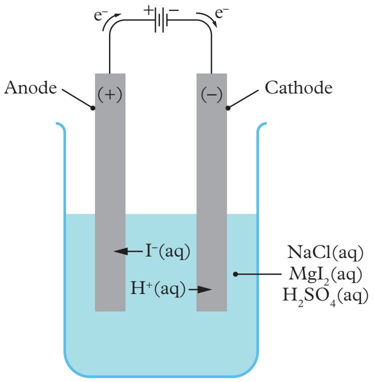
Oxidation: $2I^-(aq) \rightarrow I_2(g) + 2e^-$

Reduction: $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$

Overall: $2I^-(aq) + Co^{2+}(aq) \rightarrow Co(s) + I_2(g)$

At least 0.82 V is required.

d



Inert electrodes are made of platinum or graphite.

Strongest oxidising agent: $H^+(aq)$

Strongest reducing agent: $I^-(aq)$

Oxidation: $2I^-(aq) \rightarrow I_2(g) + 2e^-$

Reduction: $2H^+(aq) + 2e^- \rightarrow H_2(g)$

Overall: $2H^+(aq) + 2I^-(aq) \rightarrow H_2(g) + I_2(g)$

At least 0.54 V is required.

6 The Hall–Héroult cell produces CO_2 as part of the reaction. CO_2 is a greenhouse gas that, once released, enters the atmosphere and contributes to increased global temperatures.

7 The scientist can use molten NaCl as the electrolyte, as water will act as a competing reactant and will preferentially produce H_2 and O_2.

Reduction: $Na^+(l) + e^- \rightarrow Na(l)$

Oxidation: $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

At least 4.07 V is required.

8 **a** No reaction will occur.

b The zinc fork will corrode and the silver electrode will be electroplated with the silver ions from the electrolyte. The zinc fork will undergo oxidation while the silver ions from the solution will be reduced.

c 1 Clean the fork and the silver metal electrode with emery paper to remove any products of oxidation.

2 Fill the beaker with a 1.00 M solution of silver nitrate.

3 Connect the positive terminal of the battery to the silver electrode and the negative terminal to the zinc fork.

4 Apply 4 V to the cell and run for 10 minutes.

7.3 Features of rechargeable secondary cells

7.3 CHECK YOUR LEARNING

1 Primary cells cannot be recharged but secondary cells can. Primary cells have an irreversible reaction but secondary cells have a reversible reaction. Primary cells have no fluids in them but secondary cells can be made up of aqueous cells and molten salts.

2 Recharging cannot take place once the products have fallen away.

3 Power supply instead of a galvanometer or a lightbulb; a single container instead of individual half-cells

4 **a** Nickel undergoes reduction at the cathode. Cadmium undergoes oxidation at the anode.

b $2NiO(OH)(s) + Cd(s) + 2H_2O(l) \rightarrow Cd(OH)_2(s) + 2Ni(OH)_2(s)$

c Oxidation: $Ni(OH)_2(s) + OH^-(aq) \rightarrow NiO(OH)(s) + H_2O(l) + e^-$

Reduction: $Cd(OH)_2(s) + 2e^- \rightarrow Cd(s) + 2OH^-(aq)$

d At least 1.30 V is required.

5 **a** Cathode: $V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$

Anode: $VO^{2+}(aq) + H_2O(l) \rightarrow VO_2^{\ +}(aq) + 2H^+(aq) + e^-$

b $VO^{2+}(aq) + H_2O(l) + V^{3+}(aq) \rightarrow V^{2+}(aq) + VO_2^{\ +}(aq) + 2H^+(aq)$

c At least 1.25 V is required.

d Student answers may vary but should include a mention of an acidic electrolyte.

6 **a** Anode: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Cathode: $2H^+(aq) + e^- \rightarrow H_2(g)$

Overall: $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

b 0.76 V; can be used to power small electrical devices, such as small LED lights.

c Over time, the voltage supplied by the lemon battery will decrease due to the zinc electrode corroding or the lemon juice drying up as a result of the continual loss of H^+ ions. Commercial batteries may last longer and have a higher voltage output than a single lemon battery. Many commercial batteries are also rechargeable and can be reused.

d Typically, the term *battery* is used to describe a rechargeable secondary cell. Because a lemon battery is not rechargeable, it might be more accurate to describe it as a lemon cell instead.

7.4 Producing 'green' hydrogen

7.4 SKILL DRILL

1 Primary cells, such as non-rechargeable batteries, are generally thrown away after a single use, and are therefore an example of a linear economy. On the other hand, secondary cells can be reused and are therefore an example of a circular economy. Secondary cells are therefore much more sustainable.

2 'Green' hydrogen could be a better alternative to meeting society's energy needs because it contains more energy per mass compared to other sources such as natural gas and methane. It is also produced more sustainably, aligning with green chemistry principles and sustainable development goals.

3 Design for energy efficiency, prevention of wastes, use of renewable feedstocks

7.4 CHECK YOUR LEARNING

1 Hydrogen is available in limited supply in nature. Hydrogen also offers higher energy content compared to current energy sources such as natural gas and methane.

2 Green hydrogen is hydrogen that is produced in a more sustainable manner compared to the SMC method. It would be produced via electrolysis of water to produce hydrogen and oxygen gas rather than hydrogen and carbon dioxide gas.

3 q(hydrogen) = 141 kJ

q(methane) = 55.6 kJ

q(natural gas) = 54.0 kJ

4 Student answers may vary but should include at least one similarity and one difference, listed in full in your obook pro.

5 Find the full labelled drawing in your obook pro.

Strongest oxidising agent: $H_2O(l)$

Strongest reducing agent: $OH^-(aq)$

Oxidation: $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$

Reduction: $2H_2O(l)+2e^- \rightarrow H_2(g)+2OH^-(aq)$

Overall: $2H_2O(l) \rightarrow O_2(g)+2H_2(g)$

At least 2.06 V is required.

6 The production of hydrogen via electrolysis is considered green because it does not release the greenhouse gas carbon dioxide into the atmosphere, unlike SMC.

7 It is the established technology. Alternatives are much more expensive.

7.5 Faraday's Laws

7.5 CHALLENGE

1 20 g (2 sig fig)

2 1.9 cm^3 (2 sig fig)

3 6.9×10^{-3} cm (2 sig fig)

7.5 CHECK YOUR LEARNING

1 Current of circuit, time circuit is running, identity of electrolyte

2 Both use the principles of redox chemistry. The processes are similar and there will be a relationship between the current and the flow of ions in the circuit.

3 1.1 g (sig fig)

4 a 2.8 g (2 sig fig)

b 1.9 M (2 sig fig)

5 0.053 g (2 sig fig)

6 0.134 h (3 sig fig)

7 +4

8 a Student answers will vary but should include information about the electrolyte, electrode materials and values for current and time. An example response is shown in your obook pro

b Student answers will vary.

Chapter 7 review

MULTIPLE CHOICE

1	B	2	D	3	C	4	A	5	B
6	C	7	B	8	C	9	C	10	D

SHORT ANSWER

11 Electrons will not move through solutions in an electrolytic cell because they will take the path of least resistance, which is through the wires.

12 In galvanic cells, the electrons are 'freed' at the anode during oxidation. and move through the external circuit toward the positively charged cathode. In electrolytic cells, the power supply will force the electrons into the cathode to undergo reduction, resulting in a negative polarity.

13 In the galvanic cell, the reactants need to be separated because direct contact will result in a spontaneous reaction. In the electrolytic half-cell, separation of the reactants is not required because they will not react spontaneously.

14 To allow the flow of ions to the electrodes for the transfer of electrons to occur

15 After the end of the discharge, all reactants have been converted into products. During recharge, the discharge products become the reactants and are slowly converted into products, which are the original discharge reactants. The electrical energy supplied needs to be greater than the electrical energy produced during the recharge reaction. The products of the discharge need to be in contact with the electrodes for the reaction to occur.

16 The products of discharge will be lost from the cell. It will then be unable to undergo the reverse reaction to reform reactants in the recharge process.

17 The iron anode is a reactive electrode. It will be the species reacting at the anode to form iron ions instead of bromine gas.

18 Copper is the strongest reducing agent and will thus react at the anode to form $Cu^{2+}(aq)$. Water is the strongest oxidising agent and will thus react at the anode to form hydrogen gas (H_2) and hydroxide ions (OH^-).

19 Water is the stronger oxidising and reducing reagent in this scenario. Therefore it reacts in preference.

20 Student answers may vary but must refer to the use of non-standard conditions, because a concentration greater than 2.0 M will encourage the production of chlorine instead of oxygen.

21 a Anode: $2Cl^-(l) \rightarrow Cl_2(g)+2e^-$

Cathode: $Mg^{2+}(l)+2e^- \rightarrow Mg(l)$

b $2Cl^-(l)+Mg^{2+}(l) \rightarrow Cl_2(g)+Mg(l)$

c The sodium and calcium impurities help lower the melting point of magnesium chloride. The mixture allows the molten state to be maintained easily without requiring a large amount of heat.

d If the iron ion was used as one of the impurities, it would react at the cathode, because the iron ion is a stronger oxidising agent than magnesium, calcium or sodium.

e If the cell was aqueous, water would be present at both the anode and the cathode. Water is the stronger oxidising agent and reducing agent at the electrodes, so hydrogen gas and oxygen gas will form instead of the desired products of magnesium and chlorine gas.

22 a Anode is electrode I and cathode is electrode II.

b Anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

c They 'fall out' of the impure electrode as the copper in the anode becomes copper ions. They are both weaker reducing agents than copper, so they are unable to form the ions.

d Molten copper(II) sulfate electrolyte would still produce the desired copper solid at the cathode and the impure copper electrode will still oxidise to produce the copper ions required in the solution.

23 a In discharge:

Anode: $Ni(OH)_2(s)+OH^-(aq) \rightarrow Ni(OH)_3(s)+e^-$

Cathode: $Fe(OH)_2(s)+2e^- \rightarrow Fe(s)+2OH^-(aq)$

In recharge:

Anode: $Fe(s)+2OH^-(aq) \rightarrow Fe(OH)_2(s)+2e^-$

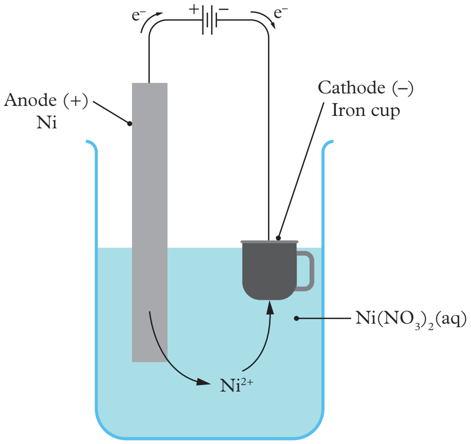
Cathode: $Ni(OH)_3(s)+e^- \rightarrow Ni(OH)_2(s)+OH^-(aq)$

b In discharge: $2Ni(OH)_2(s)+Fe(OH)_2(s) \rightarrow 2Ni(OH)_3(s)+Fe(s)$

In recharge: $2Ni(OH)_3(s)+Fe(s) \rightarrow 2Ni(OH)_2(s)+Fe(OH)_2(s)$

c Alkaline

24



Strongest oxidising agent: $Ni^{2+}(aq)$

Strongest reducing agent: $Ni(s)$

Oxidation: $Ni(s) \rightarrow Ni^{2+}(aq)+2e^-$

Reduction: $Ni^{2+}(aq)+2e^- \rightarrow Ni(s)$

Any amount of voltage will start the reaction.

25 Student answers will vary depending on the cells designed, but example responses are shown. Full labelled drawings can be found in your obook pro.

a Strongest oxidising agent: $H_2O(l)$

Strongest reducing agent: $H_2O(l)$

Oxidation: $2H_2O(l) \rightarrow O_2(g)+4H^+(aq) + 4e^-$

Reduction: $2H_2O(l)+2e^- \rightarrow H_2(g)+2OH^-(aq)$

Overall cell equation: $2H_2O(l) \rightarrow O_2(g) + 2H_2(g)$

At least 2.06 V is required.

b Strongest oxidising agent: $Ca^{2+}(l)$

Strongest reducing agent: $Cl^-(l)$

Oxidation: $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

Reduction: $Ca^{2+}(l) + 2e^- \rightarrow Ca(s)$

Overall: $2Cl^-(l) + Ca^{2+}(l) \rightarrow Ca(s) + Cl_2(g)$

At least 4.23 V is required.

c Strongest oxidising agent: $Ag^+(aq)$

Strongest agent: $Ag(s)$

Oxidation: $Ag(s) \rightarrow Ag^+(aq) + e^-$

Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Overall: not required

Any amount of voltage will start a reaction.

d Strongest oxidising agent: $Mg^{2+}(aq)$

Strongest reducing agent: $H_2O(l)$

Oxidation: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Reduction: $Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$

Overall: $2H_2O(l) + 2Mg^{2+}(aq) \rightarrow O_2(g) + 4H^+(aq) + 2Mg(s)$

At least 3.60 V is required.

26 **a** Strongest oxidising agent: $Ca^{2+}(l)$

Strongest reducing agent: $I^-(l)$

Oxidation: $2I^-(l) \rightarrow I_2(g) + 2e^-$

Reduction: $Ca^{2+}(l) + 2e^- \rightarrow Ca(s)$

Overall: $2I^-(l) + Ca^{2+}(l) \rightarrow Ca^{2+}(l) + I_2(g)$

At least 3.41 V is required.

b Strongest oxidising agent: $Cd^{2+}(aq)$

Strongest reducing agent: $Br^-(aq)$

Oxidation: $2Br^-(aq) \rightarrow Br_2(g) + 2e^-$

Reduction: $Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$

Overall: $2Br^-(aq) + Cd^{2+}(aq) \rightarrow Br_2(g) + Cd(s)$

At least 1.49 V is required.

c Strongest oxidising agent: $Sn^{4+}(aq)$

Strongest reducing agent: $I^-(aq)$

Oxidation: $2I^-(aq) \rightarrow I_2(g) + 2e^-$

Reduction: $Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$

Overall: $2I^-(aq) + Sn^{4+}(aq) \rightarrow Sn^{2+}(aq) + I_2(g)$

At least 0.39 V is required.

27 0.946 g (3 sig fig)

28 109.49 g or 1.1×10^2 g (2 sig fig)

29 **a** 2.78 L (3 sig fig)

b 0.925 M (3 sig fig)

30 21.8 min (3 sig fig)

31 13.6 A (3 sig fig)

32 2978.37 A or 3.0×10^3 A (2 sig fig)

33 +1

34 +2

35 For the electrolysis of hair follicles, the two relevant half-equations are:

$Cl^-(aq) \rightarrow Cl_2(g) + 2e^- \quad E^0 = +1.36\ V$

$2H_2O(l) + 2e^- \rightarrow 2OH^-(aq) + H_2(g)$

$E^0 = -0.83\ V$

Because they are quite far apart on the electrochemical series, it is unlikely that a change from 25°C to 37°C will cause them to swap positions. Therefore, we would expect the electrolysis reaction to occur and the method is therefore viable.

36 **a** Oxidation: $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

Reduction: $Na^+(l) + 2e^- \rightarrow Na(s)$

Overall: $2Cl^-(l) + Na^+(l) \rightarrow Na(s) + Cl_2(g)$

b Oxidation: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

Reduction: $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

Overall: $2Cl^-(aq) + 2H_2O(l) \rightarrow H_2(g) + 2OH^-(aq) + Cl_2(g)$

c Oxidation: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Reduction: $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

Overall: $2H_2O(l) \rightarrow O_2(g) + 2H_2(g)$

37 Student mind maps will vary, but should use the terms provided.

38 Student answers will vary, but an example summary is provided in your obook pro.

39 **a** Oxidation: $Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2e^-$

Reduction: $HgO(s) + H_2O(l) + 2e^- \rightarrow Hg(l) + 2OH^-(aq)$

b Overall: $Zn(s) + HgO(s) \rightarrow ZnO(s) + Hg(l)$

c They produced liquid mercury, which is a highly toxic metal.

d Overexposure to mercury will lead to mercury poisoning, which is fatal.

e Silver oxide batteries

f Oxidation: $Zn(s) + 2OH^-(aq) \rightarrow Zn(OH)_2(s) + 2e^-$

Reduction: $Ag_2O(s) + 2H^+(aq) + 2e^- \rightarrow 2Ag(s) + H_2O(l)$

g Overall: $Zn(s) + Ag_2O(s) + H_2O(l) \rightarrow Zn(OH)_2(s) + 2Ag(s)$

h The products of the silver oxide battery are zinc hydroxide and silver. Neither product is as toxic as mercury is to humans.

Unit 3 Area of Study 2 Checkpoint

MULTIPLE CHOICE

1	D	**2**	D	**3**	B	**4**	A	**5**	C
6	C	**7**	A	**8**	B	**9**	A	**10**	D

SHORT ANSWER

1 **a** If the volume is doubled, the concentration of everything is halved. The system will partially oppose this change by moving to the side with more particles. Because both reactants and products have the same number of particles, there will be no net forward or reverse reaction and no effect on percentage yield.

b If the temperature is increased in an exothermic reaction, the system will partially oppose the change by moving to a net backward reaction to store excess energy in reactant bonds and decrease the percentage yield of HBr.

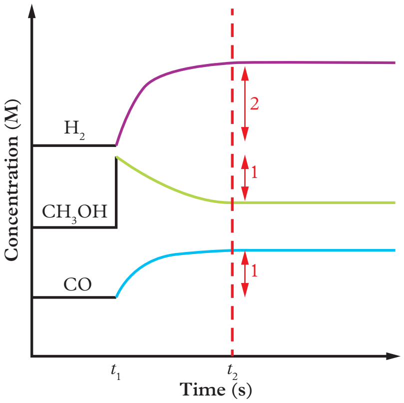
c The addition of Br_2 will initially darken the reaction mixture. The system partially opposes this change by driving net forward, resulting in a lighter colour.

2 **a** 1.5 M (2 sig fig)

b As $Q > K$, the reaction must shift net backward to reach equilibrium.

c If the volume is doubled, the concentration of everything is halved and the colour becomes lighter. The system will partially oppose this change by moving to the side with more particles, resulting in a net forward reaction. This will result in a lighter colour, because more light-yellow chlorine gas will be generated.

3 **a**



b

i $K = \dfrac{[CH_3\ OH]}{[CO][H_2]^2}$

ii 245 M^{-2}

c Rate: An increase in temperature will increase the frequency and proportion of successful collisions and therefore the rate of the reaction.

Yield: An increase in temperature in an exothermic system will result in a net backward reaction to store excess energy in reactant bonds. This will decrease the yield of methanol.

4 **a** $H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

b $H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

c →

d Sunlight is used to split water molecules into oxygen and hydrogen gases – no chemical waste is produced and the process can be repeated without emitting greenhouse gases into the environment.

5 **a** $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

b negative / –

c Na(l) is a strong reducing agent and $Cl_2(g)$ is a strong oxidising agent. If they are not separated, they will react spontaneously and violently to form sodium chloride instead of pure sodium metal.

d 1236 kg or 1.24 t (3 sig fig)

6 **a** The surface area of the rhubarb

b The time to oxidise the oxalic acid and therefore the rate of the reaction

c An increase in the surface area will increase the frequency and proportion of successful collisions and therefore the rate of the reaction.

d The precision cannot be commented on, because the experiment wasn't repeated.

e There has been no effort made to measure or control the temperature. Although not outlined, the concentration has not been reported, so you cannot be sure that it was the same for all experiments.

f The person conducting the experiment cannot be sure which variable has resulted in the change to the dependent variable.

Unit 3 review

MULTIPLE CHOICE

1	C	**2**	A	**3**	A	**4**	A	**5**	B
6	B	**7**	C	**8**	B	**9**	A	**10**	D

SHORT ANSWER

1 **a** Negative

b Neutral

c An insulating material, e.g. rubber or plastic, which helps to prevent loss of energy as heat

d Carbon/graphite

e Electrolyte paste that facilitates transport of ions between the anode and cathode

2 **a** 3.17 (2 sig fig)

b 2852 kJ mol^{-1}

3 **a** $C + 3D \rightleftharpoons 2A + B$

b *K* is large, which means that there is more product than reactant at equilibrium. This means that there is a net forward reaction.

c 1.34×10^{-6} (3 sig fig)

d Decrease

e When the temperature is increased, the equilibrium will favour the endothermic reaction, which would be the reverse reaction. Hence, more reactants will form and result in a decrease in *K*.

f 28.6 M^{-1} or 3×10^1 M^{-1} (1 sig fig)

4 **a** Cathode: $Au^+(aq) + e^- \rightarrow Au(s)$

Anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

Overall: $2Au^+(aq) + 2Cl^-(aq) \rightarrow 2Au(s) + Cl_2(aq)$

b 0.32 V

c $Au^{3+}(aq)$

d 1.68 V

e Lower, since some of the current is being used to power the bulb

5 **a** Cathode: $PbSO_4(s) + 2e^- \rightarrow Pb(s) + SO_4^{2-}(aq)$

Anode: $PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^-$

b Student answers will vary. An example response can be found in your obook pro.

6 The combustion of carbohydrates releases less energy than when the same amount of fatty acid is combusted. If more fats are consumed, there will be much more energy available to the body, so a diet high in fats is more fattening than a carbohydrate-rich diet.

7 373.1 g or 3.7×10^2 g (2 sig fig)

Chapter 8: Structure, nomenclature and properties of organic compounds

GROUNDWORK

8A Covalent bonding is chemical bonding that involves sharing of one or more pairs of valence electrons between atoms to achieve a complete valence shell.

8B Structural formula, semi-structural (condensed) formula, skeletal structure.

8C Student answers may vary, but should include the following points:

1 Determine the longest chain of carbon atoms – this will be your 'main chain'. Use the correct prefix based on the total number of carbon atoms.

2 Determine which end is nearest to an alkyl branch, functional group, double or triple bond – this is your 'starting point'. The starting point will be the carbon closest to the highest priority group.

3 Number the carbon atoms from the 'starting point' to the terminal carbon atom.

4 Identify and name any alkyl group branches.

5 Identify any functional groups present and determine whether to use prefixes or suffixes depending on their type and position.

6 When two or more branches or functional groups occur on the same or different carbon atom, the number of the carbon atom is indicated for each branch. Alkyl branches and halogens have the same priority and are named in alphabetical order.

7 When two or more identical branches or functional groups occur on different carbon atoms, the counting prefixes 'di-', 'tri-' and 'tetra-' are used.

8.1 Carbon and organic compounds

8.1 SKILL DRILL

1 Student answers will vary, but should include: a step-by-step method written in passive voice, and materials, including rubber bands.

2 Student answers may vary, but example risks include: eye injuries from the rubber bands snapping, hand injuries from pulling rubber bands apart, cuts from cutting any rubber bands.

3 The distance between the rubber bands (bond length) decreases as more rubber bands are added (bond strength).

4 Student answers may vary, but example limitations include: the rubber band elastic energy wears off with time, which does not happen with bond strength.

8.1 CHECK YOUR LEARNING

1 Organic compounds are chemical compounds that have carbon–hydrogen bonds.

2 Student answers may vary, but examples include: a carbon atom can form single, double or triple covalent bonds with other carbon atoms, and a carbon atom can form bonds with up to four carbon atoms.

3 Bond strength is the measure of how strongly atoms are joined to each other. This strength can be measured in bond energy, which is the amount of energy required to break the covalent bonds.

4 Saturated molecules have the maximum number of hydrogen atoms attached to all carbon atoms. Unsaturated molecules do not have the maximum number of hydrogen atoms attached to all carbon atoms.

5 Structural isomers are molecules with the same molecular formula, but are arranged in different ways. Their different arrangements can create unique molecules with their own physical and chemical properties, which contributes to the diversity of organic compounds.

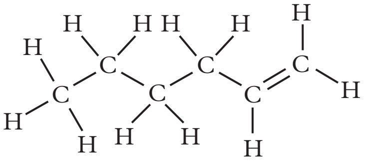
6 Chain isomers occur when molecules have the same molecular formula, but a different arrangement or branching of the main carbon chain. Positional isomers occur when molecules have the same molecular formula, but a double bond or functional group is attached to a different position along the main chain.

7 The order is incorrect. The correct order arranged from highest to lowest bond length is: C–Br > C–C > C–O > C=N

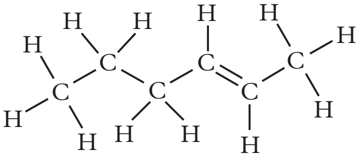
8 C=O > C=N > C–F > C–H > C–C

9 **a** DoU = 0
  b DoU = 1
  c DoU = 1

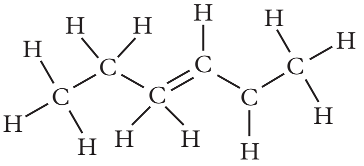
10 There are 13 non-cyclic structural isomers of C_6H_{12}. Three are shown here. The names and structures of the remaining 10 can be found in your obook pro.



Hex-1-ene



Hex-2-ene



Hex-3-ene

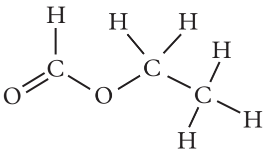
11 Atomic radius and bond length have a directly proportional relationship. As the atoms get bigger, the bond length also increases, because there is more distance between the two nuclei.

8.2 Functional groups

8.2 REAL-WORLD CHEMISTRY

1 Ethanol and methanoic acid

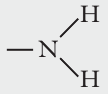
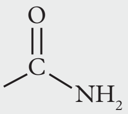
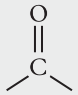
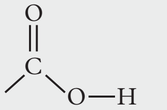
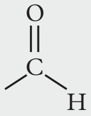
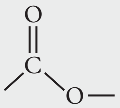
2



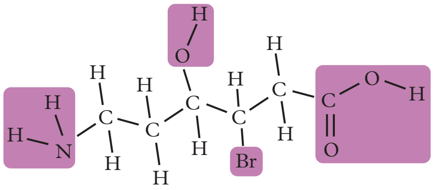
3 Common names are also simpler to read and remember as they are often shorter, and do not contain digits or hyphens.

8.2 CHECK YOUR LEARNING

1

Name of the homologous series	Structural formula of the functional group
Alkane	C–C
Primary amine	
Primary amide	
Primary alcohol	—O—H
Ketone	
Carboxylic acid	
Alkene	C=C
Aldehyde	
Ester	
Haloalkane	–F, –Cl, –Br, –I

2 **a**



b Amino group, hydroxyl group, halogen (bromo) group, carboxyl group

3 **a** Semi-structural formula
  b Skeletal structure
  c Structural formula

4 Benzene does not follow the same general formula as alkenes. The electrons in benzene are also delocalised, whereas the electrons in alkenes are associated with distinct bonds. This gives benzene different chemical and physical properties to alkenes.

5 Cyclohexane has a cyclic structure compared to straight-chained alkanes. This causes cyclohexane to have fewer than the maximum number of hydrogen atoms, but cyclohexane is considered saturated. Also, cyclohexane has a different general formula (CnH_{2n}) from other non-cyclic alkanes (CnH_{2n+2}).

6 Student answers will vary, but an example response is: Primary amines have an NH_2 group attached to the alkyl or R group. Primary amides have a carbon atom with a double-bonded oxygen atom and an NH_2 group attached to the alkyl or R group.

7 Student answers will vary but should show that primary alcohols have a carbon atom attached to one hydroxyl group and one alkyl group; secondary alcohols have a carbon atom bonded to one hydroxyl and two alkyl groups; tertiary alcohols have a carbon atom bonded to one hydroxyl and three alkyl groups.

8 Student answers will vary but should show that a ketone has a carbonyl group (C=O) attached to two different carbon atoms within the alkyl chain, whereas an aldehyde has a carbonyl group (C=O) as the terminal carbon atom in the alkyl chain.

9 The table can be found in your obook pro.

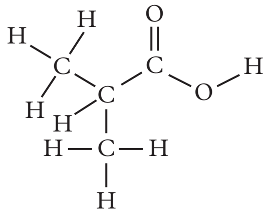
10 Alcohol ($C_nH_{2n+2}O$), aldehyde ($C_nH_{2n}O$), amide ($C_nH_{2n+1}NO$), amine ($C_nH_{2n+3}N$), carboxylic acid ($C_nH_{2n}O_2$), ester ($C_nH_{2n}O_2$), haloalkane ($C_nH_{2n+1}X$), ketone ($C_nH_{2n}O$)

8.3 IUPAC systematic naming

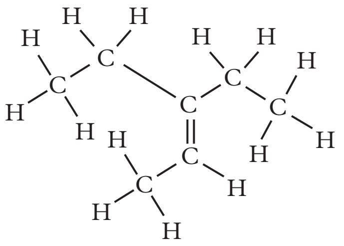
8.3 CHECK YOUR LEARNING

1 **a** I Structural formula
  II Skeletal structure
  III Semi-structural formula
  IV Skeletal structure
  b I 4-methylhept-1-ene
  II 2-aminobutanoic acid
  III ethyl butanoate
  IV 3-methylpent-1,4-diene

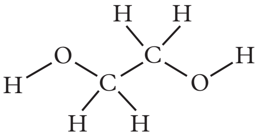
2 **a**



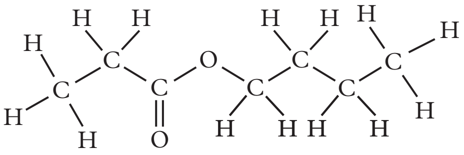
b



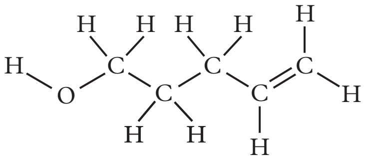
c



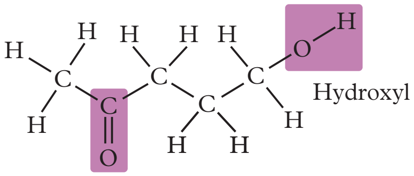
d



e



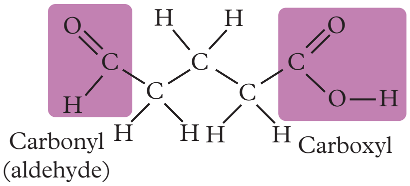
3 **a**



Carbonyl (ketone)

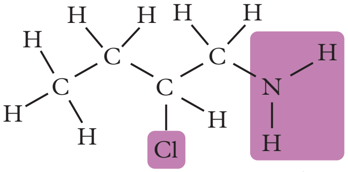
Highest priority group is the carbonyl (ketone), so the suffix must be –one.

b



Highest priority group is the carboxyl group, so the suffix must be –oic acid.

c

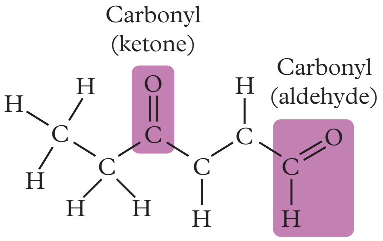


Halogen (chloro)

Amino

Highest priority group is the amino group, so the suffix must be –amine.

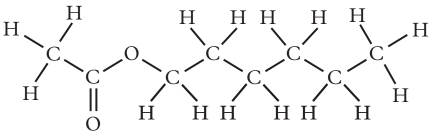
d



Highest priority group is the carbonyl group (aldehyde), so the suffix must be –al.

4 a: 5-hydroxypentan-2-one
b: 5-formylpentanoic acid
c: 2-chlorobutan-1-amine
d: 4-oxohexan-1-al

5



Hexyl ethanoate

6 The hydroxyl functional group takes priority over the amino functional group. Therefore, the first carbon atom is assigned to the end of the chain that is closest to the hydroxyl functional group so that the number associated with the hydroxyl functional group in the name is the smallest. 4-aminobutan-2-ol is the correct name for the molecule.

7 Many organic molecules have multiple functional groups. Functional group priority is important to give consistent and standardised names for these molecules. Otherwise, there would be multiple acceptable names, which would make the naming system less effective.

8.4 Physical properties of organic compounds

8.4 REAL-WORLD CHEMISTRY

1 Viscosity is the measure of the resistance of flow of a substance.

2 Frogs use their tongues to snatch insects from the air in under a second. This quick, powerful movement reduces the viscosity of the saliva into a thin water-like liquid that is quickly able to coat and trap the insect. The saliva then increases in viscosity as the tongue reels back in the frog's mouth, making it more difficult for the insect to escape.

3 Student answers may vary, but examples include: applying enough force, such as by centrifuging the saliva sample to reduce viscosity or applying heat to the saliva sample to reduce viscosity.

8.4 CHECK YOUR LEARNING

1 **a** Dipole–dipole attractions, dispersion forces
  b Hydrogen bonding, dispersion forces
  c Dipole–dipole attractions, dispersion forces
  d Hydrogen bonding, dispersion forces
  e Hydrogen bonding, dispersion forces

Chloromethane and methanal would have the lowest boiling points because of the presence of weaker dipole–dipole attractions. The presence of stronger hydrogen bonding would result in higher boiling points of methanamine, methanol and methanoic acid.

2 The greater number and strength of intermolecular bonding present between molecules, the higher their boiling point. This is because the more bonds there are, the more energy it will take to break them and change the liquid into a gas.

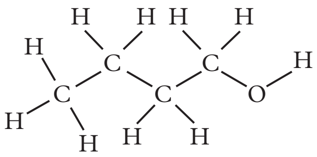
3 The melting point is the temperature at which a solid is converted into a liquid, whereas the boiling point is the temperature at which a liquid becomes a gas. The trends for melting points and boiling points of organic molecules are the same.

4 The hydroxyl groups in tertiary alcohols are more crowded than the hydroxyl groups in primary or secondary alcohols. This crowding of the hydroxyl group makes it more difficult for hydrogen bonding to occur, leading to less heat energy being required to separate the tertiary alcohol molecules (lower boiling point).

5 Propene < 2-methylbutane < pentane < hex-1-ene < octane

6 Pentan-2-one < pentan-3-one < pentanal < 1-chloropentane < pentan-2-ol < pentan-1-ol < pentanoic acid < pentanamide

7 **a** **i**



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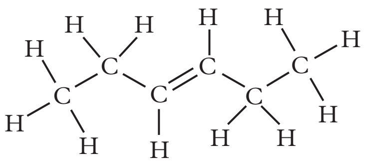
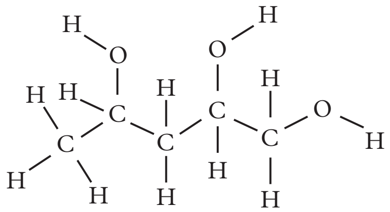


iii



b **i** Hydrogen bonding, dispersion forces

ii Dispersion forces only

iii Hydrogen bonding, dispersion forces (note: there are more opportunities for hydrogen bonding in pentan-1,2,4-triol than butan-1-ol)

c **i** Pentan-1,2,4-triol > butan-1-ol > hex-3-ene

ii Pentan-1,2,4-triol > butan-1-ol > hex-3-ene

8 Both butan-1-amine and propanoic acid can form hydrogen bonding between molecules. However, the strength of attraction between the carboxyl groups in propanoic acid is stronger than that between the amino groups in butan-1-amine because nitrogen is less electronegative than oxygen. Propanoic acid can also form dimers. More energy is therefore required to break the increased dispersion forces and then the hydrogen bonding between propanoic acid molecules compared to butan-1-amine molecules. The more energy that is required to break the intermolecular forces, the higher the boiling point.

Chapter 8 review

MULTIPLE CHOICE

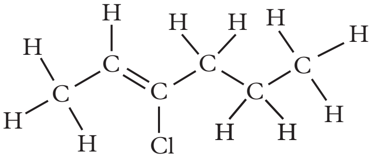
1	C	**2**	B	**3**	A	**4**	A	**5**	B
6	A	**7**	C	**8**	B	**9**	A	**10**	C

SHORT ANSWER

11 **a** Ester

b Propyl ethanoate

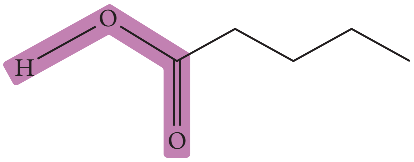
12 **a**



b This structure can have chain isomers (different lengths of the main chain and branches) and positional isomers (having the double bond and/or the chloro group in another location).

c The strongest intermolecular force that 3-chlorohex-2-ene and its isomers can make with water would be dipole–dipole interactions, because it cannot hydrogen-bond. 3-chlorohex-2-ene and its isomers will therefore have low solubility in water. Isomers identified will vary, but examples include: unbranched isomers like 2-chlorohex-2-ene and 3-chlorohex-3-ene, or a branched isomer like 2-chloro-3-methylpent-2-ene.

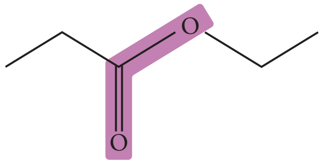
13 I



Functional group(s): carboxyl group

IUPAC name: pentanoic acid

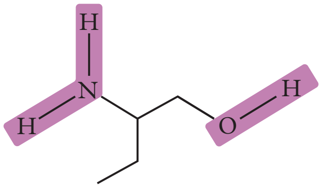
II



Functional group(s): ester group

IUPAC name: ethyl propanoate

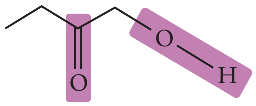
III



Functional group(s): amino group, hydroxyl group

IUPAC name: 2-amino-butan-1-ol

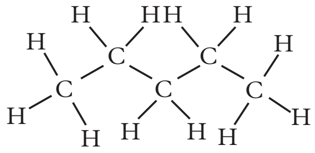
IV



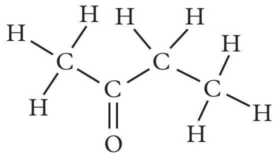
Functional group(s): carbonyl group (ketone), hydroxyl group

IUPAC name: 4-hydroxybutan-3-one

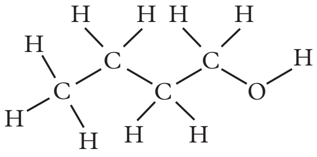
14 **a**



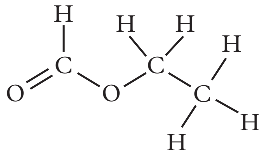
Pentane



Butanone



Butanol



Ethyl methanoate

b Pentane: dispersion forces

Butanone: dipole–dipole attractions, dispersion forces

Butanol: hydrogen bonding, dispersion forces

Ethyl methanoate: dipole–dipole attractions, dispersion forces

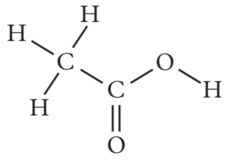
c II

d Ethyl methanoate can form dispersion and dipole–dipole attractions between its molecules, which means that more heat energy is required to separate its molecules compared to pentane, which only has dispersion forces. Butanone, however, has stronger dipole–dipole attractions between its molecules compared to ethyl methanoate due to less crowding around the functional group that participates in dipole–dipole attractions. Although butanone has a smaller molar mass and therefore weaker dispersion forces than ethyl methanoate, this decrease is less than the increase of strength in dipole–dipole interactions. Therefore, more heat energy is needed to separate the molecules of butanone than those of ethyl methanoate.

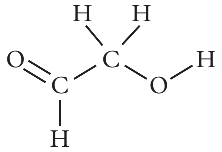
15 A double bond can be a part of a functional group (e.g. within a carboxyl group or carbonyl group) or can be a stand-alone functional group between carbon atoms (alkene group). It is considered a functional group because it is a site of comparatively high reactivity, allowing it to undergo specific chemical reactions.

16 **a** $C_2H_4O_2$

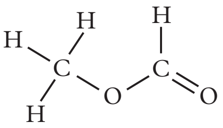
b



Ethanoic acid

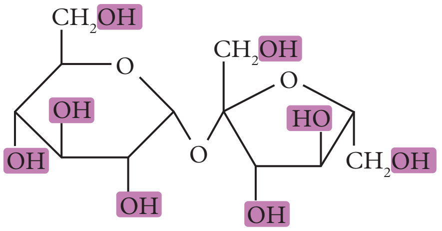


2-hydroxyethan-1-al



Methyl methanoate

17 **a** Eight hydroxyl groups



b Dispersion forces, dipole–dipole attractions, hydrogen bonding

c Maple syrup contains sucrose molecules, which form a range of intermolecular forces. This causes them to have difficulties sliding past each other, which causes the high viscosity of maple syrup.

d The intermolecular forces will start to weaken or break. This causes the viscosity of maple syrup to decrease.

18 **a** Six carbon atoms

b Two methyl alkyl groups

c 3,4-dimethylhexane

d $CH_3CH_2CH(CH_3)CH(CH_3)CH_2CH_3$

19 The correct name for this compound is pent-1-en-1-ol. The hydroxyl functional group should be added as a suffix instead of a prefix, because it is the highest priority group in the compound.

20 Ethanol contains C–C, C–H, C–O and O–H covalent bonds. Looking at bond length, the C–H bonds are the shortest, followed by C–O and C–C. This suggests that the C–H bonds are the strongest. From this chapter, we have not learnt specifically about the bond strengths of O–H bonds, but we can predict that O–H bonds are the strongest because the difference in electronegativity is greater than the C–H bond.

21 The electronegativity difference is larger in a C–F bond than a C–N bond. This larger electronegativity difference will cause the C–F bond energy to be larger and therefore the bond will be harder to break than a C–N bond.

22 Both primary and secondary alcohols contain a hydroxyl (–OH) functional group. They differ in the position of the hydroxyl group. A primary alcohol has a carbon atom with one alkyl group and one hydroxyl group. A secondary alcohol has a carbon atom which has two alkyl groups and one hydroxyl group.

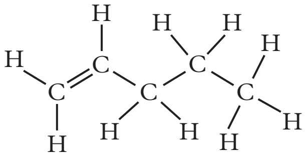
23 3,4-dimethylhepane

24 The boiling point of hex-1-ene is slightly lower than hexane because the double bond stops molecules from packing together tightly compared to the saturated chain. The dispersion forces will not be as strong between hex-1-ene molecules.

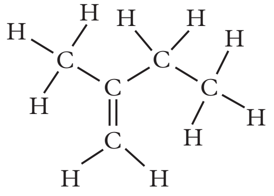
25 **a** Butanoic acid and hexan-1-ol

b Hexyl butanoate

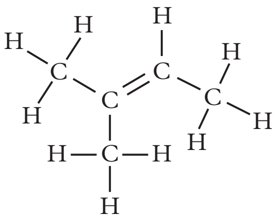
26 Three possible isomers are shown. Find the rest in your obook pro.



Pent-1-ene



2-methylbut-1-ene



2-methylbut-2-ene

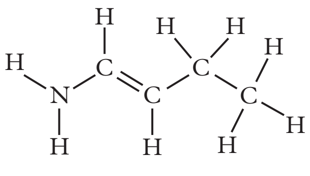
27 hexan-2-ol, 2-methylbutane, 4-methylhexan-3-ol, pent-3-en-2-amine, 3-methylbutanal

28 **a** D < C < F < E < B < A

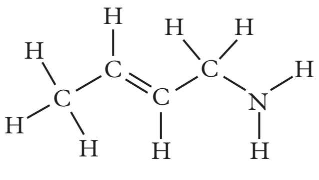
b A – propanoic acid
B – 2-methylbutan-1-ol
C – 3-methylpentane
D – 3-methylpent-2-ene
E – pentan-3-ol
F – pent-3-one

29 A – polar, high solubility in water
B – polar, high solubility in water
C – non-polar, low solubility in water
D – non-polar, low solubility in water
E – polar, high solubility in water;
F – polar, high solubility in water

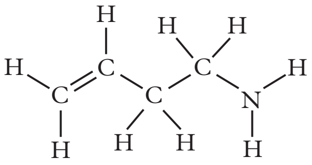
30 Molecular formula: C_4H_9N Three possible isomers are shown. Find the rest in your obook pro.



But-1-en-1-amine



But-2-en-1-amine



But-3-en-1-amine

31 $CH_3CH_2COCH_2CH_3$
Structure: carbonyl group (ketone); boiling point: 102°C
$CH_3CH_2CH_2CH_2COOH$
Structure: carboxyl group; boiling point: 186°C
$CH_3C(CH_3)_2COOH$
Structure: carboxyl group, alkyl branches; boiling point: 163.7°C

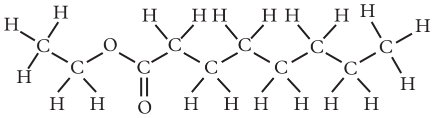
32 Both structures have the correct naming for a six-carbon-atom chain with single bonds (hexane). The hydroxyl functional group has priority over the amino functional group. Therefore, the suffix is based on the hydroxyl group (-ol) and the prefix is based on the amino group (amino-). The difference with these names is the numbering. 5-aminohexan-3-ol is correct, as the lower number is given to the hydroxyl group, which has priority over the amino group.

33 The first half of the statement is correct in that longer bonds require less energy to break. This lower bond energy leads to a less stable bond. However, the second half of the explanation is incorrect, as the C–I bond length is larger than the C–H bond length. Therefore, the C–I bond will be less stable than C–H.

34 **a** ethanol and octanoic acid

b ethyl octanoate

c



35 **a** The melting point increases along the homologous series as the alcohols become larger. More electrons are available to form dipole moments and therefore more dispersion forces can form.

b The boiling point increases along the homologous series as the alcohols become larger. More electrons are available to form dipole moments and therefore more dispersion forces can form.

c Pentan-1-ol will be a liquid at room temperature because room temperature lies between its melting (–78°C) and boiling points (138°C).

d Butan-1-ol: 118°C
Hexan-1-ol: 157°C

e Approximately –34°C

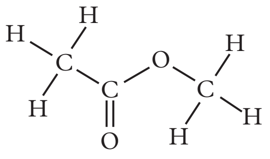
36 The functional group is a carboxyl group, which must be on the terminal carbon atom. Following IUPAC rules, the carbon atom within the carboxyl group is assigned as the first carbon atom in the chain.

37 The boiling point of an ester increases as the size of the ester molecule increases. This trend is due to the increasing number of dispersion forces with larger molecules. Overcoming more dispersion forces requires more heat energy, which leads to a higher boiling point.

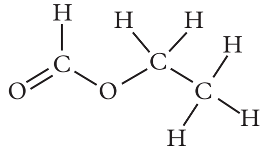
38 Correct IUPAC name: 3,3-dimethylheptan-1-ol. The student incorrectly identified the longest chain. This caused an error with the types and locations of the branches in the IUPAC name.

39 They are positional isomers where the two halogen (bromo) groups are in different locations in each structure. The names of the compounds are: 1,2-dibromobutane, 1,3-dibromobutane, 2,3-dibromobutane and 2,2-dibromobutane.

40 **a**



Methyl ethanoate



Ethyl methanoate

b The two compounds are structural isomers of each other: they have the same molecular formula but different structures. More specifically, these compounds are positional isomers, where the ester functional group is in a different location in the compounds.

41 The viscosity of the compounds from highest to lowest is C > B > A.

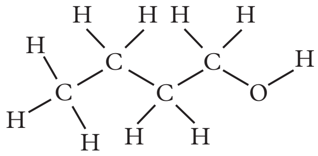
42 **a** A larger difference in electronegativity between atoms forming a covalent bond leads to a higher bond energy, which increases the bond strength.

b A higher amount of bond energy leads to greater bond stability.

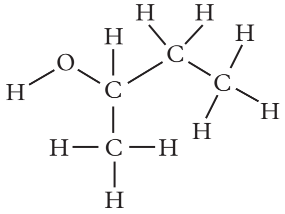
43 **a** Student answers will vary but should include a step-by-step method, written in a passive voice, and a list of materials.

b IV: the type of organic compound (X or Y); DV: boiling point

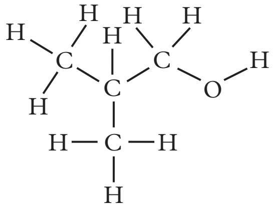
c



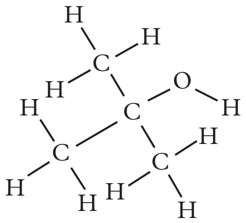
Butanol



Butan-2-ol



2-methylpropan-1-ol



2-methylpropan-2-ol

d Secondary alcohols have a lower boiling point than similar-sized primary alcohols because the hydroxyl functional group is crowded. This reduces the strength of the hydrogen bonds that can form between molecules. Branched molecules have a lower boiling point than unbranched molecules, because the molecules cannot pack together as closely, which leads to weaker dispersion forces.

Using these rules, the alcohols are ordered as follows: secondary branched alcohol < secondary unbranched alcohol < primary branched alcohol < primary unbranched alcohol.

e Student answers will vary but should mention how the absolute boiling points could be determined, as well as how this information can be used to determine the two unknown molecules.

44 **a** The incorrect main chain was identified, leading to the wrong alkyl group names and main chain numbering. The correct name is 2,3-dimethylpentane.

b The name provided is correct.

c Incorrect numbering, because carboxyl groups have the highest priority and contain the first carbon atom in the chain. The correct name is 5-hydroxypentanoic acid.

d Incorrect functional group prefix and suffix used, because hydroxyl groups have higher priority than amino groups. The correct name is 4-aminobutan-1-ol.

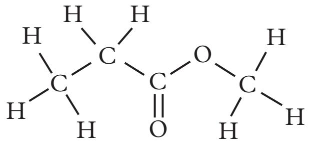
e Incorrect numbering and functional group prefix and suffix used, because aldehydes have a higher priority than hydroxyl groups. Numbering should start from the carbon atom in the aldehyde group. The correct name is 3-hydroxypentan-1-al.

f Incorrect main chain was identified, leading to the wrong alkyl group names and main chain numbering. The correct name is 1-amino-3-methylbutan-2-one.

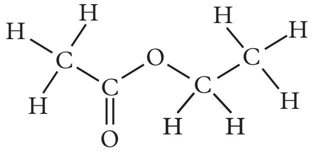
g Missing counting prefix and location of double bond not indicated. The correct name is 2,2,3-trichlorohex-3-ene or 2,2,3-trichlorohex-4-ene or 2,2,3-trichlorohex-5-ene (as double bond location is unknown).

h Incorrect suffix used. The correct name is propanamine.

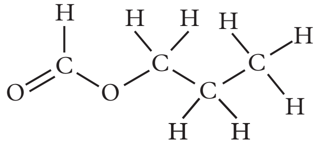
45 **a**



Methyl propanoate



Ethyl ethanoate



Propyl methanoate

b Propyl methanoate: 81°C; methyl propanoate: 79.8°C; ethyl ethanoate: 77.1°C

c These esters can all form dispersion forces and dipole–dipole attractions between their molecules. The dipole–dipole attraction is due to the polarity caused by the ester functional group. The closer this ester functional group is to one side of the molecule, the more accessible it is to form dipole–dipole attractions with other molecules. The level of accessibility of this ester functional group from lowest to highest is ethyl ethanoate < methyl propanoate < propyl methanoate, which corresponds to their boiling points.

46 A primary alcohol has a carbon bonded to one hydroxyl group and one alkyl group. A primary amine has a carbon bonded to one amino group and one alkyl group.

47 Student answers will vary. You can find a sample response in your obook pro.

Chapter 9: Categories, properties and reactions of organic compounds

GROUNDWORK

9A Student answers will vary, but may include: hydroxyl (–OH), carboxyl (–COOH), halogens (F, Cl, Br or I), carbonyl (–C=O) and amine (–NH_2).

9B A catalyst is a chemical that is added to a reaction to reduce its activation energy. This will increase the rate of reaction. A catalyst does not get consumed in a reaction and can be reused.

9C The principles of green chemistry aim to minimise the impact of chemistry products or processes on the environment.

9.1 Organic reactions

9.1 SKILL DRILL

1 IV: The types of oils used (coconut, vegetable and olive)

DV: Hardness, smell, gentleness of soap on hands

2 Hydrolysis reactions break up larger molecules into smaller molecules.

3 Student answers will vary, but may include specifying: the type of tray/moulds used to set the soap, temperature to keep soap while setting, the steps to test hardness and the steps to test gentleness of soap – using multiple people/hands.

4 The students decided that no particular oil produced the best soap because no one particular soap scored ideal results on the three measures used (hardness, smell and gentleness on hands).

5 Student answers will vary, but may include: 'What other oils make the best soap (using other oils)?', 'What bases make the best soap (trying other strong bases instead of sodium hydroxide)?' or 'What oil combinations make the best soap (mixing the oils in different combinations/ratios)?'

9.1 CHECK YOUR LEARNING

1 **a** A reaction where two or more molecules react to form one new molecule

b A reaction where an atom on one molecule is replaced by another from another molecule, so if you start with two reactants, you will end with two new products

c A reaction that results in a chemical species losing electrons

2 **a** $KMnO_4$ or $K_2Cr_2O_7$

b Heat/UV

c H_3PO_4

d H_2SO_4

e $K_2Cr_2O_7$

3 **a** Oxidation

b Substitution

c Addition

d Condensation

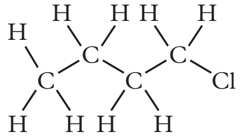
e Oxidation

4 Reaction 1: But-1-ene+hydrochloric acid

a Addition

b There are none.

c

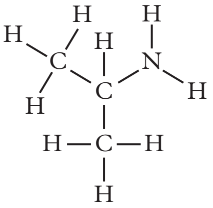


Reaction 2: 2-chloropropane + ammonia

a Substitution

b There are none.

c



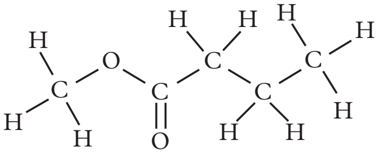
H—H

Reaction 3: Methanol+butanoic acid

a Condensation

b H_2SO_4

c

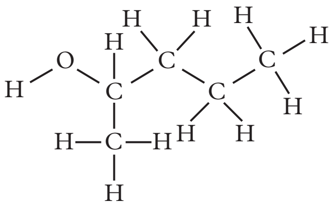


Reaction 4: Pent-2-ene+water

a Addition

b H_3PO_4

c



5 Find the full pathways with structural formulas in your obook pro.

a Box 1: H_2O (water)

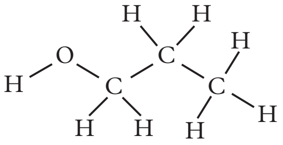
Box 2: HCl

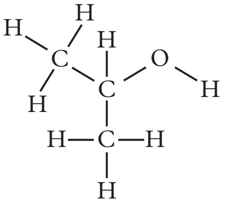
Box 3: CH_3CH_2OH (ethanol)

b Box 1: $CH_3CH_2CH_2$ (propene)

Box 2: Ni

6 **a**





b The isomers can form from the oxidation of propene using H_3PO_4 and a high temperature (300°C). Find the full pathways with structural formulas in your obook pro.

c The primary alcohol, propan-1-ol, oxidises to form propanal (an aldehyde), whereas the secondary alcohol, propan-2-ol, oxidises to form propanone (a ketone).

7 Both reactions often involve water but the aim of hydrolysis reactions is to break down large molecules, whereas the aim of condensation reactions is to form larger molecules by joining smaller ones.

8 Find the full pathways with structural formulas in your obook pro.

a Forms 2,3-dibromobutane

b Forms butan-2-ol

c Forms 2-chlorobutane

9.2 Organic reactions of large biomolecules

9.2 CHECK YOUR LEARNING

1 The condensation reaction results in a polypeptide with four amino acids and three glycosidic bonds, and it produces three water molecules. Find the full pathway with structural formulas in your obook pro.

2 Find the full pathways with structural formulas in your obook pro.

a Ester links form from the reaction between the –COOH group on the fatty acid and the –OH group on the glycerol molecule.

b Glycosidic (ether) links form between the –OH group of the glucose and the CH_2OH of the sucrose, producing a disaccharide and water.

c An amide/peptide bond forms between the C=O and the –NH, producing a peptide and water.

d A glycosidic link forms between OH groups, producing cellulose and water.

3 Converting a triglyceride into a biodiesel molecule uses the addition of three alcohol molecules. A small alcohol like methanol is added and then the solution is warmed with a hydroxide catalyst (either potassium hydroxide or sodium hydroxide). This process is known as transesterification and it produces three biodiesel molecules and one glycerol molecule, as shown in your obook pro.

4 **a** A: peptide bond
B: amino acid residue
C: carboxyl group

b Glutamic acid, isoleucine, tyrosine

5 **a** Similarities: Both are polysaccharides of glucose.
Differences: The polysaccharides are composed of different glucose units (alpha in amylose and beta in cellulose), which changes the bonding orientation of glycosidic links.

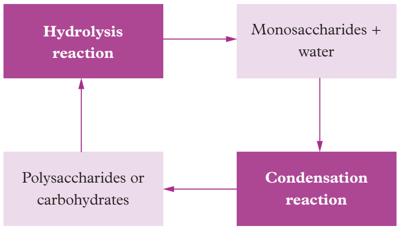
b Cellulose is hydrolysed to form beta glucose molecules. Amylose is hydrolysed to form alpha glucose molecules. The full pathways with structural formulas can be found in your obook pro.

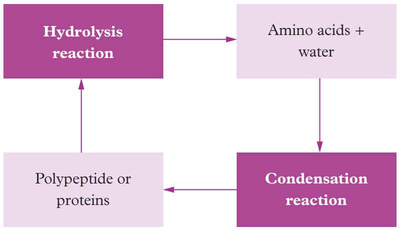
c Humans do not have the enzymes to catalyse the reaction needed to break down beta glycosidic links.

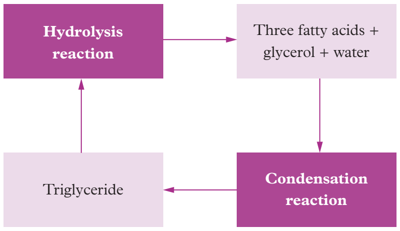
6 Similarities: Both can occur inside living organisms, involve organic molecules, involve water molecules and can happen to all macromolecules (protein, fats and carbohydrates).

Differences: Hydrolysis is the breakdown of larger molecules while condensation is the building of larger molecules from smaller ones.

7







9.3 Atom economy and percentage yield

9.3 CHECK YOUR LEARNING

1 The percentage atom economy is the comparison between the molar mass of desired product with the molar mass of all reactants based on the chemical equation. The percentage yield is the comparison between the actual yield of a desired product with the theoretical yield of a desired product.

2 39%

3 **a** Reaction 1: addition
Reaction 2: substitution

b Reaction 1: HCl
Reaction 2: NaOH

c Reaction 1: 100%
Reaction 2: 63%

4 Reaction 1: 58.5%
Reaction 2: 41.5%

5 Reaction 1 has a higher percentage atom economy, which will better adhere to the green chemistry principle due to the decreased weight of reactants required and higher efficiency.

6 Ethene → ethanol (addition reaction with water)

Hexane → chlorohexane (substitution reaction with chlorine gas) → hexanol (substitution reaction with water) → hexanoic acid (oxidation with oxygen gas)

Ethanol + hexanoic acid → ethyl hexanoate

7 Student answers will vary but should include the addition of hydrogen gas to ethene, addition of a hydrogen halide to ethene, addition of water to ethene, and the addition of ethene to ethene.

Students should calculate the atom economy for each reaction using the correct formula.

8 Student answers will vary.

9.4 Sustainability of the production of chemicals

9.4 CHECK YOUR LEARNING

1 Catalyst selection in the past was based on maximising percentage yield with very little regard for waste production. By following the green chemistry principles, companies now select catalysts based on atom economy, with a focus on waste reduction.

2 Renewable feedstocks are raw materials that are replenished as fast as, or faster than, they are consumed, and they are used as the starting materials to create a desired chemical. This is different from using fossil fuels as feedstocks, as fossil fuels are formed over millions of years and cannot be replenished as fast as they are consumed.

3 Bioethanol and biobutanol energy content is lower than that of petrol. More fuel is required to yield the same amount of energy. The burning of bioethanol and biobutanol still produces greenhouse gases, which contributes to air pollution. The viscosity of biofuels is also different from that of petrol, which means that they cannot be used interchangeably with petrol.

4 *Bio-based polymers* are plastics made from biomass instead of fossil fuel feedstock. *Biodegradable* plastics refers to the ability of such plastics to break down over time – not all biodegradable plastics come from renewable sources. Similarly, not all bio-based polymers are biodegradable.

5 The production of bio-based polymers is very similar to the production of conventional fossil fuels and only differs by one step. While the production of biopolyethylene requires a conversion from biomass into ethene, fossil fuel production does not have this step.

An advantage of using bio-based polymers is that the growth of biomass can offset large amounts of carbon dioxide through photosynthesis. However, the production of bio-based polymers requires intensive agricultural practices to produce the crops.

6 Use of renewable feedstocks: Raw materials or feedstocks should be made from renewable (mainly plant-based) materials, rather than from fossil fuels, whenever practicable. This is aligned with responsible consumption and production, concentrating on the limited global supply of fossil fuel reserves.

Catalysis: Catalysts should be selected to generate the same desired product(s) with less waste and using less energy and fewer reagents in reaction processes/pathways. This is aligned with responsible consumption and production, concentrating on minimising wastage.

Designing safer chemicals: Chemical products should be designed to achieve their intended function while minimising toxicity. This is aligned with responsible consumption and production, concentrating on minimising the use and production of toxic matter.

7 Student answers will vary. A sample response for the production of bioethanol can be found in your obook pro.

8 *Bio-based polymers* should be used instead of *bioplastics* to avoid confusion with biodegradable plastics.

Chapter 9 review

MULTIPLE CHOICE

1 B **2** C **3** C **4** A **5** A
6 D **7** D **8** D **9** C **10** A

SHORT ANSWER

11 a $KMnO_4$
b Nickel
c Heat/UV
d H_3PO_4; 300°C

12 a Oxidation
b Addition
c Substitution
d Addition

13 a The full pathway with structural formulas can be found in your obook pro.
b Bioethanol production involves the use of enzymes as catalysts and produces carbon dioxide as by-product. Ethanol from fossil fuel production utilises H_3PO_4 as catalyst.
c Bioethanol and ethanol from fossil fuels are chemically and molecularly the same. The only difference is their origin.

14 Because the activation energy from methane straight into methanol is too high, alternative chemical pathways are required to form methanol from methane. A possible pathway is:
Methane → chloromethane (by substitution reaction with chlorine gas) → methanol (by substitution reaction with water)

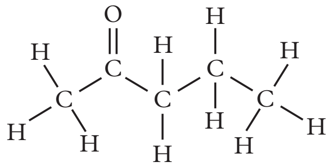
15 The hydroxyl group in pentan-2-ol is located on the middle carbon atom of the hydrocarbon chain. This makes pentan-2-ol a secondary alcohol. When secondary alcohols oxidise, ketones are formed rather than carboxylic acids. Carboxylic acids are formed from the oxidation of primary alcohols.

16 Student answers will vary but may include: minimising reaction steps, minimising waste by-products, minimising products or reactants loss or using a catalyst.

17 Addition reactions will have high atom economies (100%) because all reactants are incorporated into the one product that is required. Condensation reactions, in general, will also have a high atom economy because the only by-product produced is water, which has a relatively low molecular weight.

18 Atom economy is an important consideration when designing reaction pathways to reduce the amount of wastage.

19 a The reaction forms pentan-2-one. The full pathway with structural formulas can be found in your obook pro.



b i The full pathway with structural formulas can be found in your obook pro.
ii 52% (2 sig fig)
iii No, no alternative pathways with higher atom economy can be used.
c The full pathway with structural formulas can be found in your obook pro.

20 1,2-difluoro-3-methylbutane; Reactants will include 3-methylbut-1-ene and fluorine.

21 a Compound K can be 4-chloro-2-methylhexane or 2-methylhex-3-ene, and Compound M is likely to be 5-methylhexan-3-one.
b If a haloalkane was used, a catalyst will be used. If it was an addition reaction from an alkene, then H_3PO_4 will be used.
c 5-methylhexan-3-ol
d Addition/substitution and oxidation
e Reaction 1 from chloroalkane: 76%
Reaction 1 from alkene: 100%
Reaction 2: 98%

22 Amylopectin has more branches and hydrolyses faster. This is because the amylopectin particles cannot be packed as closely and have more exposed bonds that can be hydrolysed faster than those of amylose.

23 Pathway 1: overall 52%
Pathway 2: overall 85%
Pathway 2 is more efficient because its first-reaction atom economy was 100%, whereas pathway 1's first-reaction atom economy was only 77%. Both pathways' second-reaction atom economies are the same.

24 a A: glycosidic, B: ester, C: peptide
b A: disaccharide, B: triglyceride, C: polypeptide
c A: monosaccharides, B: glycerol + 3×fatty acids, C: amino acids
d The full pathway with structural formulas can be found in your obook pro.

25 a Propene can be reacted with water in an addition reaction to form two different isomers. As seen in the reaction in your obook pro, both isomers can then undergo oxidation reactions.
b Propan-1-ol; propan-2-ol; propanoic acid
c To propanoic acid: 81% (2 sig fig)
To propanone: 97% (2 sig fig)

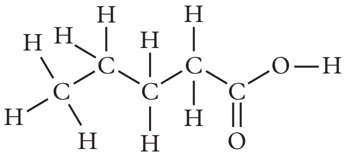
26 a Substitution
b 48.4% (3 sig fig)
c HF is the limiting reagent because there is a 1 : 2 ratio and the amount of CCl_4 is more than half of HF.
d 57.1% (3 sig fig)

27 a 94% (2 sig fig)
b The change in catalyst will not change the atom economy, because the same reaction will occur; however, it may change the percent yield.

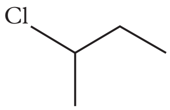
28 a Glucose
b Both are branched polymers of glucose; however, glycogen has more branching than amylopectin.
c Water molecules react with glycosidic bonds that hold the amylopectin molecule together, breaking up the molecule into smaller units.
d Glucose molecules will come together in a condensation reaction to form glycogen while one water molecule gets released with every new glycosidic bond formed.

29 a Glycerol + linolenic acid + linoleic acid + oleic acid
b It is more likely to be a plant triglyceride because the acids are unsaturated.
c Liquid – the unsaturated fatty acids create kinks, which decreases the melting point, meaning less heat energy is required to change from solid to liquid state.
d The full pathways with structural formulas can be found in your obook pro.

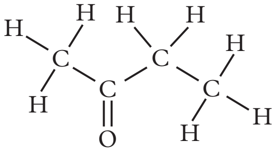
30 a H_3PO_4
b



c 2-chlorobutane

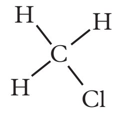


d

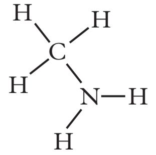


e Oxidation

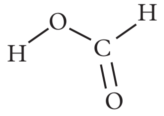
31 **a** Q: chloromethane



R: methylamine

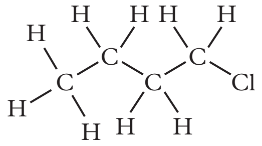


S: methanoic acid

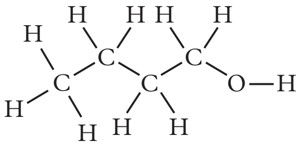


b Student answers may vary depending on the position in which chlorine substitutes to, but an example response is:

N: 1-chlorobutane



P: butan-1-ol



c B: H_2O
C: $KMnO_4/\ H^+$

d Butyl methanoate

e Left: substitution, substitution
Right: substitution, oxidation
Bottom: esterification/condensation

f 45% (2 sig fig)

32 Student answers will vary but may include any combination of the advantages and disadvantages listed in your obook pro.

33 Butane → substitution reaction with Cl_2 → chlorobutane → substitution reaction → butanol → oxidation → butanoic acid

Butanol + butanoic acid → butyl butanoate

34 **a** Yes; chloroethene could be synthesised from bioethanol through a series of reactions.

b Ethanol collected from biomass → dehydration → ethene → addition → 1,2-dichloroethane → thermal cracking → vinyl chloride monomer

35 Student answers will vary but may include: using a catalyst to increase the rate of reaction rather than to increase pressure/temperature, choosing reaction pathways aimed at increasing atom economy, reducing waste or reducing toxic waste, or sourcing renewable feedstock when available.

36 **a** Ethane → reaction with Cl_2 → chloroethane → reaction with water → ethanol

But-1-ene → addition reaction with water → butan-1-ol → oxidation → butanoic acid

Ethanol + butanoic acid → ethyl butanoate

b Student answers will vary but should calculate the percentage atom economy and refer to the green chemistry principles to justify their choices.

Unit 4 Area of Study 1 Checkpoint

MULTIPLE CHOICE

1	B	**2**	C	**3**	D	**4**	A	**5**	A
6	D	**7**	D	**8**	B	**9**	C	**10**	B

SHORT ANSWER

1 **a** $CH_3CH_2CH_2CH_2Cl$

b Substitution reaction

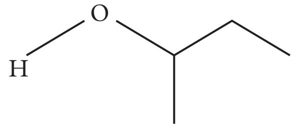
c $C_4H_{10}O$

d $KMnO_4/H^+$ or $K_2Cr_2O_7/H^+$ (or the ions)

e Butanoic acid

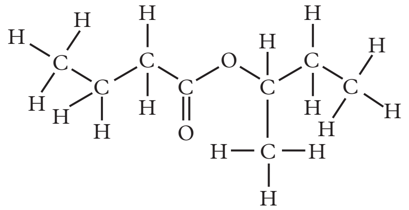
f H_3PO_4 and 300°C (also requires H_2O as a reactant)

g



h Butan-2-one

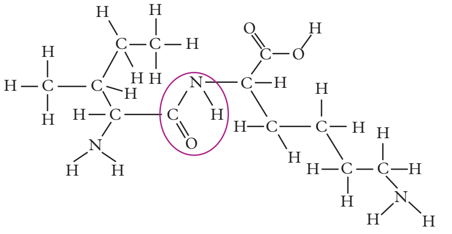
i



2 **a** Molecule X is a triglyceride; molecule Y is a protein; molecule Z is a carbohydrate (polysaccharide).

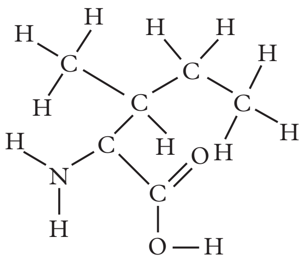
b **i** Molecule Y

ii



iii Peptide link or amide

iv



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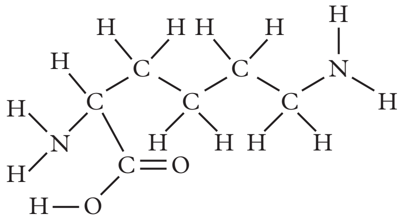


c The fatty acid contains 18 carbon atoms, with one terminal carbon being part of a carboxyl (COOH) group. The full structure can be found in your obook pro.

d **i** Molecule Z

ii Ether or glycosidic bond

iii Hydrolysis reaction

3 **a** The full pathway with structural formulas can be found in your obook pro.

b The full pathway with structural formulas can be found in your obook pro.

c $25.3\,kJ\,g^{-1}$ (3 sig fig)

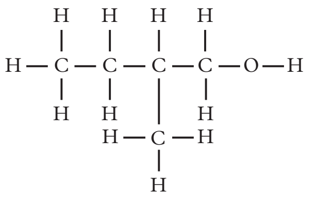
d Peanut oil has the lowest energy content according to the student's results, followed by canola oil. Olive oil has the highest energy content.

e All three oils come from plants that can be grown and replanted. Therefore, they can be considered renewable feedstocks.

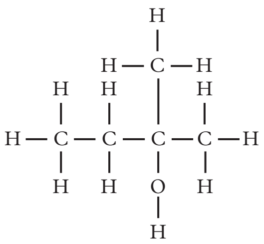
4 **a** Hydrogen bonding

b Pentan-1-ol has hydrogen bonding, whereas pentanal has only dipole–dipole attractions between molecules. Hydrogen bonding is a much stronger intermolecular force than dipole–dipole attractions, so it will take more energy to break the intermolecular forces between pentan-1-ol molecules. Therefore, pentan-1-ol has a higher boiling point than pentanal (in which it will take less energy to break the dipole–dipole attractions between molecules).

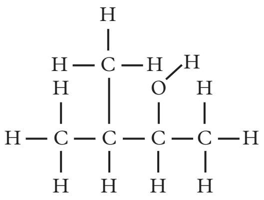
c



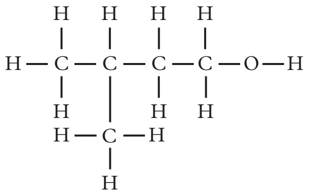
2-methylbutan-1-ol



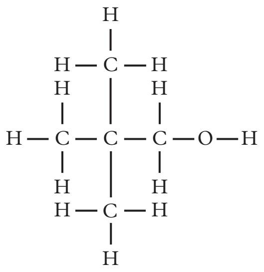
2-methylbutan-2-ol



3-methylbutan-2-ol



3-methylbutan-1-ol



2,2-dimethylpropan-1-ol

d A branched isomer of pentan-1-ol will have a lower boiling point than pentan-1-ol. Although the strongest intermolecular force is still hydrogen bonding, the branching will not allow the molecules to pack together as tightly as would a chain without branches. This means that dispersion forces between branched molecules will not be as strong as those between straight-chain molecules with the same molar mass.

e The trend in viscosity will be the same as the trend in boiling point. The stronger the intermolecular forces between molecules, the higher the viscosity. Pentanoic acid will have the highest viscosity, followed by (in decreasing order) pentan-1-ol and pentanal. The lowest will be pentane.

Chapter 10: Laboratory analysis of organic compounds

GROUNDWORK

10A An oxidising agent is the reactant that undergoes reduction in a redox reaction.

10B The melting point is the temperature at which a substance changes its state from solid to liquid. The boiling point is the temperature at which a substance changes its state from liquid to gas. The boiling point is higher than the melting point.

10C Molecules of saturated organic compounds only contain single bonds between their atoms. Unsaturated organic compounds contain double and/or triple bonds.

10D A titration is a form of volumetric analysis in which a standard solution is added to a solution of unknown concentration, until the reactants have reacted in equal mole ratios. The steps involved are outlined in your obook pro.

10.1 Qualitative chemical tests

10.1 SKILL DRILL

1 A Saturated; reduces permanganate; not reactive to acid; not reactive to base; no carboxyl group

B Saturated; cannot reduce permanganate; not reactive to acid; reactive to base; carboxylic acid

C Unsaturated; reduces permanganate; not reactive to acid; not reactive to base; no carboxyl group

D Unsaturated; cannot reduce permanganate; not reactive to acid; reactive to base; carboxylic acid

2 Qualitative analysis/data is observable but not measurable. Quantitative analysis/data is measurable.

3 1 Testing for saturation

2 Testing for reactivity with permanganate/reducing power

3 Testing for reactivity with HCl/alkalinity

4 Testing for reactivity with NaOH/acidity

5 Testing for ability to form esters/presence of carboxyl functional group

10.1 CHALLENGE

Adding an indicator to check the pH: If the pH is low, then the indicator will turn to its acidic colour. Reacting with a base such as sodium hydrogen carbonate: carbon dioxide gas will be produced (gas bubbles). Reacting with ethanol and sulfuric acid as catalyst: if the chemical produced has a fruity smell, it indicates an ester.

10.1 CHECK YOUR LEARNING

1 Iodine test (with starch indicator): 1 double bond + I_2 (coloured)→ iodohydrocarbon (colourless); e.g. $C_2H_4+I_2 \rightarrow C_2H_4I_2$

Bromine test: 1 double bond + Br_2 (coloured) → bromohydrocarbon (colourless); e.g. $C_2H_4+Br_2 \rightarrow C_2H_4Br_2$

2 Qualitative tests are only observable, not measurable, and enable the chemist to find out further properties of the sample. Quantitative tests are measurable and enable the chemist to tell how much there is of a substance or its property. Student examples may vary but could include: iodine test to find the iodine value (quantitative)/degree of saturation or bromine test to determine the presence of double bonds.

3 Oxidising ethanol with dichromate:

$$3CH_3CH_2OH(aq)+2Cr_2O_7^{2-}(aq)+16H^+(aq) \rightarrow 3CH_3COOH(aq)+4Cr^{3+}(aq)+11H_2O(l)$$

Oxidising ethanol with permanganate:

$$5CH_3CH_2OH(aq)+2MnO_4^-(aq)+6H^+(aq) \rightarrow 5CH_3COOH(aq)+2Mn^{2+}(aq)+8H_2O(l)$$

4 **a** 1 Trial mixing the sample with water to determine polarity – but-2-ene is non-polar and should not mix with water; e.g.

$$CH_3CHCHCH_3(l)+H_2O(l) \rightarrow C_4H_8(l) + H_2O(l)$$

2 Testing for degree of unsaturation with iodine; e.g.

$$CH_3CHCHCH_3+I_2 \rightarrow CH_3CHICHICH_3(l)$$

b Size of molecule (length of carbon chain) and position of double bond

5 **a** Testing with ethanol and sulfuric acid as catalyst to check for formation of fruity ester; e.g.

$$CH_3CH(OH)CH_2CH_2COOH(aq)+CH_3CH_2OH(aq) \rightarrow CH_3CH(OH)CH_2CH_2COOCH_2CH_3(aq)$$

Testing with a pH indicator: carboxylic acids ionise in water to form the hydronium ion; e.g.

$$CH_3COOH(aq)+H_2O(l) \rightleftharpoons CH_3COO^-(aq)+H_3O^+(l)$$

   b Size of molecule (length of carbon chain) and checking for multiple hydroxyl functional groups

6. Testing with ethanol and sulfuric acid as a catalyst to check for formation of a fruity ester:
   $CH_3CH(OH)CH_2CH_2COOH(aq)+CH_3CH_2OH(aq) \rightarrow CH_3CH(OH)CH_2CH_2COOCH_2CH_3(aq)$
   Testing with magnesium ribbon to see if hydrogen gas is produced; because ethanoic acid is a weak acid, it will not completely dissociate, resulting in the following reaction:
   $CH_3COOH(aq)+Mg(s) \rightleftharpoons CH_3COO^-(aq)+Mg^+(aq)+H_2(g)$
   The hydrogen ion on the ethanoic acid is displaced by the magnesium, forming hydrogen gas. This reaction can also be represented as: $CH_3COOH(aq)+Mg(s) \rightleftharpoons CH_3COOMg(aq)+H_2(g)$

10.2 Purity of consumer products

10.2 SKILL DRILL

1. Qualitative analysis, because it is an observation of whether it was maintained above or below 100°C.
2. The samples used in the experiment have fairly close boiling points, which makes it difficult to separate the different components well, resulting in contaminated samples.
3. Slower increase in temperature; use smaller sample amounts each time; repeat the distilling process to refine.

10.2 CHECK YOUR LEARNING

1. Kinetic energy is increased when a sample is being heated. As the kinetic energy increases, the particles have more movement and more energy to overcome the intermolecular forces between them, resulting in the particles melting or boiling.
2. At the melting point, there is enough kinetic energy to partially break the intermolecular forces/solid lattice. When the temperature is further increased to the boiling point, the temperature is hot enough to vaporise the sample (change its state from liquid to gaseous). At the boiling point, there is enough kinetic energy to completely overcome intermolecular forces, freeing the molecules to move around in a gaseous state.
3. Pure chemicals have a specific melting point, while impure chemicals have a lower and broader range at which they melt. This is due to the impurities disrupting the solid crystalline structure that holds a solid substance together. This lowers the amount of energy and therefore heat required to disrupt the intermolecular forces between molecules or ions.
4. Simple and fractional distillation are separation techniques whereby parts of the sample are vaporised to rise through a condenser and then turn back into a liquid. Fractional distillation can be used to separate components with similar boiling points, whereas simple distillation can accurately separate components that have very different boiling points.
5. $CH_3OH(aq)$, $CH_3CH_2OH(aq)$, $CH_3CH_2CH_2OH(aq)$,
6. $CH_3CH_2COOH(aq)$, $CH_3CH_2OH(aq)$, $CH_3CH_3(aq)$
7. **a** Sucrose molecules have a fairly large molecular weight and many hydroxyl functional groups. They can form very strong intermolecular forces of all types.
   b As sucrose melts, the intermolecular forces holding the molecules together in a solid lattice weaken. They are still strong enough to hold the molecules reasonably close together in a liquid state, but not strong enough to hold them in the rigid solid lattice structure.
8. Products that are pure will have specific and narrow melting and boiling points. If products contain contaminants, they may melt and boil within a broader range. Melting points may also be lower due to contaminants disrupting the crystalline solid structure. Student examples will vary, but should describe one consumer product. Discussion should highlight the importance of verifying components in the product and assessing its purity.
9. The claim is correct. Melting point determination is quantitative because it generates a value or range of values for the melting point of a substance. It is also qualitative because the experimental melting point can be compared against literature values to confirm the identity of the substance.

10.3 Iodine titrations to test unsaturation

10.3 CHECK YOUR LEARNING

1. A primary standard is a solute that is dissolved into a solvent to create a standard solution that has a highly accurate known concentration.
2. The equivalence point is the theoretical point where all the reactants are used up in the reaction and there are no reactants in excess. The end point is the visual cue that the equivalence point has been reached.
3. The degree of unsaturation is a measurement of the number of double bonds in a substance.
4. Addition reaction
5. **a** 0.05 M (1 sig fig)
   b 17 g (2 sig fig)
6. 4
7. 10 L (2 sig fig)
8. If the burette was wet with water, the concentration of the content inside the burette will be lower than what it was meant to be, due to the increased water. If the burette was contaminated with the remnants of a more concentrated solution, the concentration of the content inside the burette will be higher than what it was meant to be, due to the leftover particles.
9. **a** The amount of solid weighed out will greater. Therefore, the actual concentration of the standard solution will be higher than calculated/expected.
   b The actual amount of standard, in mol, used to react with the analyte will be higher than calculated/expected.
   c The actual concentration of the analyte will be higher than calculated/expected.

10.4 Redox titrations for volumetric analysis

10.4 CHALLENGE

23%(*m*/*m*) (2 sig fig)

10.4 CHECK YOUR LEARNING

1. Some of the chemicals commonly used in redox reactions change colour after undergoing redox reactions, so no additional indicator is required. For example, redox reactions involving permanganate are often self-indicating because this substance changes from purple to clear when reduced. On the other hand, redox reactions involving iodine need starch indicator added to visualise the end point.
2. Student answers will vary, but because permanganate is an oxidising agent that undergoes reduction, the organic compound selected must be a reducing agent that undergoes oxidation. Examples include alcohols and aldehydes.
3. Student answers will vary, but because oxalic acid is a reducing agent that undergoes oxidation, the organic compound selected must be an oxidising agent that undergoes reduction. Examples include carboxylic acids, ketone and aldehydes.
4. **a** $S_2O_3^{2-}(aq)+5H_2O(l) \rightarrow 2SO_4^{2-}(aq)+10H^+(aq)+8e^-$

b $C_2H_2O_4(aq) \rightarrow 2CO_2(g) + 2H^+(aq) + 2e^-$

c $CH_3CH_2OH(aq) + H_2O(l) \rightarrow CH_3COOH(aq) + 4H^+(aq) + 4e^-$

5 **a** Oxidation: $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

b Reduction: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

c Overall: $3Sn^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 3Sn^{4+} + 2Cr^{3+}(aq) + 7H_2O(l)$

d 0.228 M (3 sig fig)

6 **a** Oxidation: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

b Reduction: $H_2C_2O_4 \rightarrow 2CO_2 + 2H^+ + 2e^-$

c Overall: $5H_2C_2O_4 + 2MnO_4^- + 6H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

d 0.0365 M (3 sig fig)

e Difference: 0.0165 M or 45% – this could be due to glassware not being rinsed with the correct solutions before use, or incorrect calculation of the standard solution concentration due to weighing errors.

Chapter 10 review

MULTIPLE CHOICE

1	A	**2**	B	**3**	D	**4**	C	**5**	C
6	A	**7**	A	**8**	B	**9**	C	**10**	D

SHORT ANSWER

11 Student answers will vary, but could explain the bromine and iodine tests.

12 Student answers may vary. A sample answer explaining one test is available in your obook pro.

13 Esterification reactions can be used to determine the presence of a carboxylic acid; the following can also be used to identify the presence of an acid in general: using a pH indicator, acid–base (carbonates) reactions, or reaction with metals. A sample answer explaining one test is available in your obook pro.

14 Sodium chloride is an ionic compound whose solid lattice is held together by very strong ionic bonds that require high amounts of energy to overcome, which results in a very high melting point. Water and methane are both molecular compounds, but while water molecules are polar, methane molecules are not. Water can form all three types of intermolecular forces, giving it a higher melting point compared to many other molecular compounds. On the other hand, methane molecules are small and non-polar, which means that they can only form dispersion forces. This results in a very low melting point.

15 Substances with higher boiling points have stronger intermolecular forces because more energy is required to completely break these intermolecular forces and set the molecules free.

16 Impurities can disrupt the crystalline structure that holds a solid substance together, lowering the amount of energy and therefore the heat required to break the intermolecular forces between molecules or ions.

17 Aliquot: the volume of solution dispensed, using a pipette, into the conical flask titration

Titre: the volume of solution dispensed from the burette

18 Permanganate is a stronger oxidising agent than dichromate.

19 If the burette was wet with water, the concentration of the content inside the burette will be lower than what it was meant to be due to the increased water. This will result in the concentration of the unknown solution being calculated to be higher than it is supposed to be.

20 Leaving water in the conical flask will have no effect on the final concentration because it does not affect the amount of aliquot particles being tested.

21 Both ethanol and water molecules are able to form all three types of intermolecular bonds, which results in them having lower boiling points. However, while ethanol includes two carbon hydrogen chains (non-polar parts), water molecules contain highly polar bonds and are able to form hydrogen bonds in a tetrahedral formation. This results in water having a very high boiling point, which is higher than that of ethanol.

22 Both methane and octane are non-polar hydrocarbon molecules that are only able to form dispersion forces. The strength of dispersion forces in molecules is dependent on the molecular size/weight. Because octane is much larger than methane, it has much stronger dispersion forces. This results in octane having a much higher boiling point than methane.

23 **a** Methanol will have a higher boiling point. Methanol is able to form all three types of intermolecular bonds, whereas methane can only form dispersion forces.

b Butane will have a higher boiling point. Both methane and butane are non-polar hydrocarbon molecules that are only able to form dispersion forces. The strength of dispersion forces in molecules is dependent on the molecular size/weight. Because butane is larger than methane, it has much stronger dispersion forces.

c Ethanoic acid will have a higher boiling point. Both ethanoic acid and ethanol are able to form all three types of intermolecular bonds. However, ethanoic acid can also form two hydrogen bonds (and therefore, dimers), which increases its boiling point.

24 **a** $2Br^- \rightarrow Br_2 + 2e^-$

b $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

c $2MnO_4^-(aq) + 16H^+(aq) + 10Br^- \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Br_2$

d 22.2 mL (3 sig fig)

25 **a** $2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$

b $I_2 + 2e^- \rightarrow 2I^-$

c $2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$

d 0.0039 mol (2 sig fig)

e 0.0019 mol (2 sig fig)

f 4.9% (2 sig fig)

26 **a** $Fe \rightarrow Fe^{3+} + 3e^-$

b $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

c $3MnO_4^-(aq) + 24H^+(aq) + 5Fe \rightarrow 3Mn^{2+}(aq) + 12H_2O(l) + 5Fe^{3+}$

d 2.2 g (2 sig fig)

e 22%(*m*/*m*) (2 sig fig)

27 **a** $SO_2 + 2H_2O \rightarrow SO_4^- + 4H^+ + 3e^-$

b $I_2 + 2e^- \rightarrow 2I^-$

c $2SO_2 + 3I_2 + 4H_2O \rightarrow 2SO_4^- + 6I^- + 8H^+$

d 0.000 079 7 mol (3 sig fig)

e 0.000 053 1 mol (3 sig fig)

f 136 ppm (3 sig fig)

28 **a** $H_2O_2(aq) \rightarrow 2H^+(aq) + 2e^- + O_2$

b $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

c $Cr_2O_7^{2-}(aq) + 8H^+(aq) + 3H_2O_2(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3O_2$

d 9.7%(*m*/*v*) (2 sig fig)

29 Student answers may vary, but an example flowchart is available in your obook pro.

30 Student answers may vary, but an example flowchart is available in your obook pro.

31 Both molecules have the same types of atoms and atomic weight and are polar, which means that they have the same amount of dispersion forces and can both form dipole–dipole bonds. However, the oxygen atom of the propanol is attached to a hydrogen atom on the end. This enables hydrogen bonding as well. This results in propanol having a much higher boiling point than ethyl methyl.

32 **a** The measured titre will be higher than actual, resulting in a higher concentration being found than what it actually is.

b The measured titre will be higher than actual, resulting in a higher concentration being found than what it actually is.

c The number of reacted molecules will be lower than actual, resulting in a lower titre volume. The calculated concentration will be lower than what it actually is.

d The number of reacted molecules will be higher than actual, resulting in a higher titre volume. The calculated concentration will be higher than what it actually is.

e The concentration of the aliquot will be lower, resulting in a higher titre volume. The calculated concentration will be lower than what it actually is.

f The concentration of the titrant will be lower, resulting in a higher titre volume. The calculated concentration will be higher than what it actually is.

Chapter 11: Instrumental analysis of organic compounds

GROUNDWORK

11A A chemical compound in which one or more atoms of carbon are covalently linked to atoms of other elements.

11B Organic molecules are held together by covalent bonds.

11C Student answers will vary, but should describe techniques covered in Chapter 11, such as the bromine test to identify unsaturated compounds.

11D Qualitative analysis is used to identify observable properties or characteristics of chemical compounds. Quantitative analysis is used to determine the numeric value of certain properties, such as the concentration or the amount of a compound.

11.1 Mass spectrometry

11.1 CHECK YOUR LEARNING

1 The base peak represents the most intense or abundant fragment, which is usually created by a single split in the original ion. The molecular ion peak is the peak that occurs when the whole original molecule becomes positively charged. It is the ion of the unfragmented molecule.

2 **a** The peak at $m/z = 98$ is produced by the molecular ion containing two ^{35}Cl atoms. The peak at $m/z = 100$ is produced by the molecular ion containing one ^{35}Cl atom and one ^{37}Cl atom. The peak at $m/z = 102$ is produced by the molecular ion containing two ^{37}Cl atoms.

b 102, 100 and 98

3 **a** 86 g mol^{-1}

b 57

c $[CH_3CH_2CH_2CH_2]^+$

d C_6H_{14}

e $[CH_3CH_2CH_2]^+$

4 **a** CH_3CH_2CHO

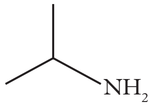
b $CH_3CH_2CH_2CH_2COOH$

5 **a** $C_3H_7NH_2$

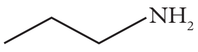
b Spectrum 1: $m/z = 44$

Spectrum 2: $m/z = 30$

c



Spectrum 1



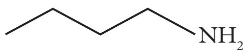
Spectrum 2

6 **a** It is the molecular ion peak produced by the presence of the carbon-13 isotope.

b 73 g mol^{-1}

c The base peak at $m/z = 44$ is caused by the loss of CH_3CH_2. 73 – 29 = 44 g mol^{-1} which corresponds to $[OCNH_2]^+$.

d



7 Yes, a mass spectrum can be used to differentiate between a ketone and an aldehyde with the same molecular formula. An aldehyde group is located at the end of the chain and a ketone group is in the middle of the molecule, so they produce different fragments.

11.2 Infrared spectroscopy

11.2 CHECK YOUR LEARNING

1 There are many functional groups in organic chemistry that can each play a significant role in the properties of different compounds and materials. IR can be used to determine the presence of functional groups and the fingerprint area can also be used to determine the identity of some molecules.

2 The peak at approximately 3500 cm^{-1} is a broad strong peak. The peak at approximately 2900 cm^{-1} is a narrow strong peak. The peak at 1700 cm^{-1} is a narrow strong peak.

3 **a** C=O

b C–H

c O–H

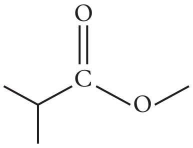
4 Amine groups.

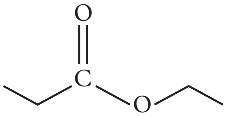
5 There is an O–H stretch with a broad peak at 3400 cm^{-1}. There is no C=O stretch, which means the compound is not a carboxylic acid.

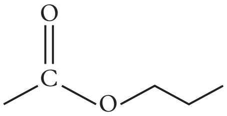
6 **a** Spectrum 1: C–H at 3000 cm^{-1}, C=O at 1750 cm^{-1}, C–O for esters or alcohols at 1250 cm^{-1}

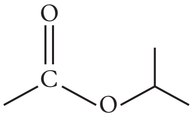
Spectrum 2: O–H at 3000 cm^{-1} due to the broad strong peak, C=O at 1750 cm^{-1}

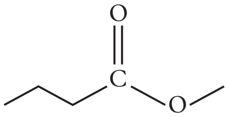
b











c Pentanoic acid (a carboxylic acid)

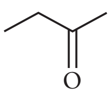
7 C–H at 3000 cm^{-1}, C=O at 1750 cm^{-1}, C–O for esters and alcohols at 1000 cm^{-1}; it does not have an O–H stretch as a broad peak at 3400 cm^{-1}, so it cannot be an alcohol. It does not have N–H at 3500 cm^{-1}, so it cannot be an amine.

11.3 Carbon-13 nuclear magnetic resonance spectroscopy

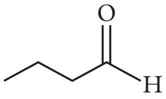
11.3 CHECK YOUR LEARNING

1 ^{13}C-NMR can be used to determine the type of environment of carbon atoms in organic compounds. ^{13}C-NMR uses the magnetic field to flip the spin of the nuclei to a high energy state, which is then returned to the original state and releases the absorbed energy. The energy is then detected, which provides information about the environment of the carbon atoms.

2 **a**



Ketone



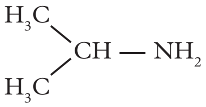
Aldehyde

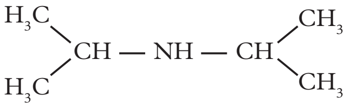
b For both the ketone and the aldehyde, four peaks will appear because there are four different environments for each molecule. The chemical shifts will be different for the two molecules because the distance between each carbon to the C=O group is different.

c If only ^{13}C-NMR is used, it would be very challenging to confirm. Despite the carbon atoms having a different chemical shift, the difference is small between aldehyde and ketones, so it is difficult to distinguish between the two molecules.

3 There are three peaks for the carbon at positions 1, 2 and 3. It is a symmetrical structure so the environment for the carbon at 1 and 5, 2 and 4 are the same. The chemical shift of the methyl groups falls in 8–25 ppm, C–OH falls in 50–90 ppm and the middle carbon falls in 20–45 ppm.

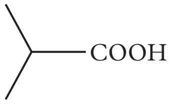
4





5 **a** It is because two carbon atoms have the same environment, so they only have one peak on the spectrum.

b



6 **a** There are two carbon environments because the two methyl groups have the same environment.

b There are three carbon environments because the structure is symmetrical. The two methyl groups have the same environment. The two $-CH_2-$ groups have the same environment. The carbon atom in the middle is another environment.

c Four environments, because each carbon atom is a different distance from the –COOH group. However, the middle two $-CH_2-$ groups would have very similar shifts.

d Five environments. Apart from the two methyl groups on the right-hand side, which have the exact same environment, the remaining four carbon atoms are all connected differently.

11.4 Hydrogen-1 nuclear magnetic resonance spectroscopy

11.4 CHECK YOUR LEARNING

1 ^{1}H-NMR can be used to determine the number of H environments and the ratio of the number of H atoms in each environment. It can also be used to determine the number of hydrogen atoms on the neighbouring carbon atom. ^{1}H-NMR can also be used to determine which functional groups are connected to a certain H or the H-associated carbon atom by chemical shift.

2 **a** Two peaks. The first peak is from the H connected to the middle carbon: integration ratio 1 as a septet (7); chemical shift is 3–4.5 ppm. The second peak is from the 6 H on the carbon atoms on the left side: integration ratio 6 as a doublet; chemical shift is 0.9–1 ppm.

b Three peaks. The first peak is the H at the carboxylic acid as a singlet; chemical shift is 9–13 ppm. The second peak is the middle hydrogen connected to the –C– as a septet, with a chemical shift of 1.5–2 ppm. The third peak is the 6 H on the side as a doublet, with a chemical shift of 1.3–1.4 ppm.

c Three peaks. The first peak is the hydrogen at the aldehyde group as a singlet, with a chemical shift of 9.4–10 ppm. The second peak is the two hydrogen atoms at $-CH_2-$ as a quartet, with a shift of 3.3–3.7 ppm. The third peak is the side 3 H as a triplet, with a shift of 0.9–1 ppm.

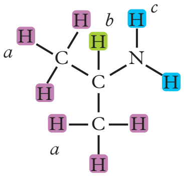
3 The *n* + 1 rule is used to determine the number of neighbouring hydrogen atoms. An environment is 'neighbouring' if it is up to three bonds away. *n* = the number of neighbouring hydrogen atoms + 1: this will equal the splitting on the neighbouring peak. Equivalent hydrogen environments won't split each other. A hydrogen in an –O–H will never cause splitting and will not be split. A hydrogen atom surrounded by two equivalent peaks will be split by all the surrounding hydrogen atoms +1.

4 From left to right: singlet, quartet, singlet, triplet and singlet

5 Peaks from left to right: A, B, C

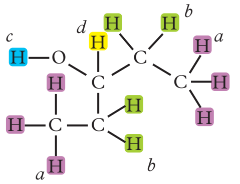
A is part of an aldehyde group with a high chemical shift. B is connected to –O– with a chemical shift of 4 ppm. It also has three neighbouring H, so it is a quartet. C is $-CH_3$ so it has a low chemical shift (1 ppm). It has two neighbouring H, so it is a triplet.

6 I Three environments



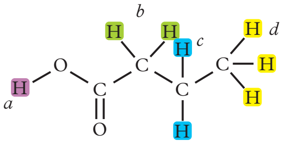
Three environments
a : *b* : *c* = 6 : 1 : 2

II Four environments



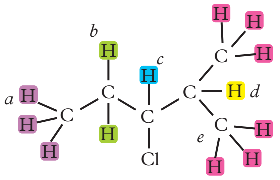
Four environments
a : *b* : *c* : *d* = 6 : 4 : 1 : 1

III Four environments



Four environments
a : *b* : *c* : *d* = 1 : 2 : 2 : 3

IV Five environments



Five environments
a : *b* : *c* : *d* : *e* = 3 : 2 : 1 : 1 : 6

11.5 Determining structures of simple organic compounds

11.5 CHECK YOUR LEARNING

1 **a** 73

b Amine peak at 3300–3500; C=O stretch for amide at around 1650

c Three peaks: one from the hydrogen atoms connected to the nitrogen atom, one from the CH_2 connected to the C=O, one from the terminal CH_3.

d Three peaks: the C=O carbon has a shift at approximately 170–190 ppm; the CH_2 carbon at 20–45 ppm; the CH_3 carbon at 8–25 ppm.

2 MS: molecular ion peak of 118. Base peak is likely to be after it loses three methyl groups, which is 73. There are some other peaks when $-CH_2CH_3$ are lost.

IR: a broad peak at 3000–3400 cm^{-1} as a result of the alcohol –OH.

^{13}C-NMR: five peaks. The carbon connected to two –OH groups would have a shift of 50–90 ppm. The –CH would have a shift of 40–60 ppm and the $-CH_2$ would have a shift of 20–45 ppm. There are two $-CH_3$ peaks with a shift of 8–25 ppm.

^{1}H-NMR: five peaks, with a peak : area ratio of 6 : 1 : 2 : 2 : 3 from left to right.

The splittings would be: doublet, septet, singlet, quartet and triplet. The expected shifts would be 0.9–1, 1.3–1.4, 1–6, 1.3–1.4 and 0.9–1.

3 **a** There should be two molecular ion peaks at 78 and 80, which correspond to the presence of chlorine-35 and chlorine-37, respectively.

b 2-chloropropane

4 **a** CH at 2500–3500 cm^{-1} and C=O of an ester at 1750 cm^{-1}

b Three hydrogen environments

c $-CH_3$ at the end of the chain.

d The presence of a $-CH_2CH_3$ isolated from a $-CH_x$ group

e Four

f The carbon at the ester group –O–C=O

g $CH_3CH_2COOCH_3$ or $CH_3COOCH_2CH_3$

h Methyl propanoate because of the chemical shift of the hydrogen atoms in the OCH_3 group (this is the most likely structure) or ethyl ethanoate

5 **a** They have the same molecular formula and will have the same molecular ion peaks.

b $CH_3CHOHCH_2CH_2CH_3$

It has five carbon environments due to the asymmetrical structure, corresponding to the five peaks in the ^{13}C-NMR spectrum. It also has a broad peak at approximately 3400 cm^{-1} in the IR spectrum corresponding to an –OH group.

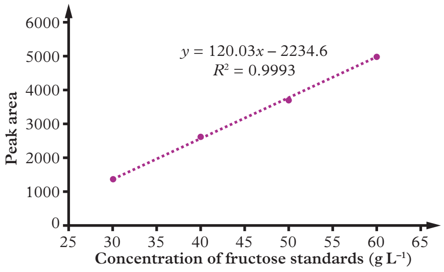
6 One type of spectroscopy can only provide a limited amount of information. The instrumental analysis techniques provide different pieces of information, which can be combined to provide a full picture of the structure of the molecule.

7 ^{1}H-NMR, because the peak area ratio, splitting and chemical shift can be used to determine the number of hydrogen environments and the neighbouring environments.

11.6 High performance liquid chromatography

11.6 SKILL DRILL

1



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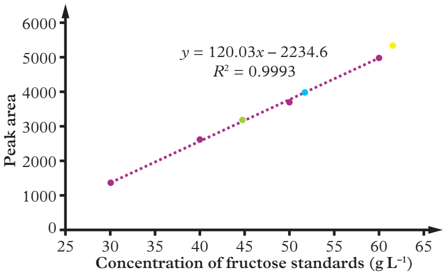


Ginger beer (green dot): 45 g L^{-1}

Lemonade (blue dot): 52 g L^{-1}

Coca-Cola (yellow dot): 61 g L^{-1}

3 The concentration of Coca-Cola is outside of the highest concentration of the standards.

4 Only one experiment was performed for each sample or standard. This could cause random errors and affect the precision of their results. Only four concentrations were used for the standard curve, and the concentration of one sample was out of the range of the standards. This could cause systematic errors and affect the accuracy of their results.

5 Increase the number of samples (at least three) to improve precision and increase the number of standards to improve accuracy. Make sure the concentration of samples is within the concentration range of the standards.

11.6 CHECK YOUR LEARNING

1 A set of known concentrations is first run through HPLC. Then, using the chromatogram from HPLC, the peak area can be calculated, which is then used to plot a standard curve of concentrations versus peak area. The peak area of the unknown samples is then plotted onto the standard curve and calculated.

2 Different conditions cause different affinities between the analyte and the stationary phase or mobile phase, so analytes might come out at different times, preventing the identification of the compound.

3 Two different substances might have similar affinities to the stationary and mobile phases, so the same retention time for both is possible.

4 **a** Pentane, methane, ethanol and butanoic acid; the mobile phase is non-polar, so the most non-polar material would come out first.

b The order would be reversed: butanoic acid, ethanol, methane and pentane. The most polar molecule would come out first because it has the highest affinity with the mobile phase. The least polar molecule would come out last.

5 **a** The polarity from high to low is Z, Y, X, W.

b The hexane is Y, because the HPLC chromatogram shows that Y has a retention time of approximately 7.1 minutes under the same conditions.

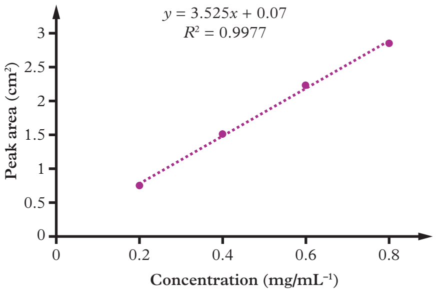
c Yes. In HPLC, as long as the run and the calibration were performed under the same conditions, the method is validated and repeatable, so the same compound would have the same retention time.

6 **a** X could be glucose; Y could be either sucrose or fructose; Z is probably taurine.

b The structures of the molecules are similar and have similar affinity with the mobile and stationary phases.

c Change the polarity of the mobile phase or stationary phase or use the reverse polarity of the phases to improve the resolution of X and Y.

7 **a**



b 0.52 mg/mL

c 195 mg (3 sig fig)

11.7 Applications of laboratory and instrumental analysis

11.7 CHECK YOUR LEARNING

1 Medical products are extremely important for our health. Impurities could be harmful by counteracting the active ingredients or causing toxic side effects.

2 Analysis of cosmetics is important to make sure that finished products are safe for consumers; that, if restricted chemicals are used, their quantities are within safe limits; that raw materials are of high quality; and that products adhere to mandatory labelling requirements.

3 Oxidation: $H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^-$

Reduction: $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

Redox titration of H_2O_2 with potassium permanganate $KMnO_4$ can be used to determine the concentration of H_2O_2 in the hand sanitiser solution. The $KMnO_4$ in the titrant is a dark purple colour;

it is reduced when added to the H_2O_2 and becomes colourless. When all the hydrogen peroxide in the aliquot is used up, the colourless solution will become and remain purple.

4 **a** 160 mg/g
  b 368 mg (3 sig fig)
  c 92.0% (3 sig fig)

5 3.0%(*m*/*v*) (2 sig fig)

6 If low-purity fuel is used over a long time, it might damage the engine and cause safety problems. This also decreases the lifetime of the vehicle.

Chapter 11 review

MULTIPLE CHOICE

1 D **2** D **3** B **4** D **5** B
6 C **7** A **8** C **9** D **10** B

SHORT ANSWER

11 **a** Sharp peak at 3300–3500 cm^{-1}
  b Medium to high sharp peak for the –OH group at 2500–3500 cm^{-1} and C=O sharp strong peak at 1680–1740 cm^{-1}
  c Broad strong peaks at 3200–3600 cm^{-1}
  d C=O sharp strong peak at 1720–1840 cm^{-1} and C–O peak at 1050–1410 cm^{-1}, but it might be hidden in the fingerprint area

12 To keep the analysis of compounds consistent across all NMR spectrometers, and allow sharing of results between chemists. TMS is chemically inert and produces only one single peak. All chemical shifts are relative to the shift of the TMS peak, which is assigned as 0.

13 **a** Chemical shift at ~175 ppm, which corresponds to the RCOOH carboxyl C or R-C=O=OR ester carbon
  b From left to right: A, B, C, D
  A, with the peak at ~70 ppm, could be an amine carbon ($R_3C–NH_2$, $R_3C–NR_2$) or an alcohol carbon ($R–CH_2–OH$) or RC≡CR).
  B, with the peak at ~32 ppm, could be $R–CH_2–R$ or $R_3–CH$.
  C, with the peak at ~22 ppm, is $R–CH_2–R$.
  D, with the peak at ~10 ppm, is $R–CH_3$.
  c From left to right: A, B, C
  A, with the peak at ~47 ppm, is $R_3–CH$.
  B, with the peak at ~27 ppm, is $R–CH_2–R$.
  C, with the peak at ~10 ppm, is $R–CH_3$.

14 The mobile phase is polar, so the most polar molecule would come out first. The most polar molecule is the one with the most hydrogen bonding forming groups and the best ability to form dipole–dipole interactions. The –COOH group is more polar than –OH because of its greater ability to form dipole–dipole interactions. Therefore, the order of the molecules coming out are: aspirin, paracetamol and ibuprofen.

15 **a** 16.5 min
  b The mobile phase is polar, so the most polar compound would come out first. The order from the most to the least polar is: A > B > C > D.
  c It provides information about the concentration of the compounds. The higher the concentration, the bigger the area.

16 The quality of cosmetics can be determined by running a sample through IR; it can be used to identify the presence of the functional group of the active ingredient. The presence of functional groups that should not be in the sample can also be detected to determine if an impurity is present.

17 There is only one C peak and one H peak because all the carbon and hydrogen atoms in the cyclohexane structure are equivalent. The chemical shift for the carbon peak would be expected in the 20–40 ppm range. The chemical shift in the ^{1}H-NMR spectrum should be in the 1–2 ppm range.

18 **a** Broad band at 3000 cm^{-1} is O–H for a carboxylic acid; C=O at 1750 cm^{-1}.
  b O–H at 3200 cm^{-1}; C–H at 3000 cm^{-1}.
  c N–H at 3500 cm^{-1}; C–H at 3000 cm^{-1}.
  d O–H at 3000 cm^{-1}; C=O at 1750 cm^{-1}.

19 Two peaks. One peak for the 6 H (two CH_3 groups) on the sides of the molecule as a triplet, and a chemical shift of 0.9–1 ppm. One peak for the 2 H in the middle of the molecule (on CH_2 group) as a septet (7), with a chemical shift of 1.3–1.4 ppm.

20 1.9%(*m*/*v*) (2 sig fig)

21 **a** Propan-1-ol
  b Propan-2-ol
  c Propanoic acid
  d Methyl ethanoate

22 $CH_3–CH_2–CH_2–NH_2$

23 **a** G
  b F or J
  c I
  d B
  e A or C
  f H
  g D

24 Isomers and their environments can be found in your obook pro.

25 **a** Paracetamol: 151
  Aspirin: 180
  Ibuprofen: 206
  b O–H at 3000 cm^{-1}; C=O at 1750 cm^{-1}
  c Ibuprofen

26 **a** The structural isomers of octanol can be found in your obook pro.
  b No. IR is useful for identifying functional groups. All the isomers have the same functional group, which is –OH, so it would be difficult to distinguish between them.
  c The hydrogen and carbon environments for the isomers of octanol can be found in your obook pro.

27 **a** C–H at 3000 cm^{-1}; C=O at 1750 cm^{-1}; the organic compound has C=O but without O–H or N–H. It is likely a ketone or aldehyde.
  b IR provides information about the type of functional groups present but does not provide information about how many of those functional groups there are. It also does not provide the molecular mass or the ways that functional groups are connected.

28 **a** Spectrum 1: –OH stretch at 3000 cm^{-1}; C=O stretch at 1750 cm^{-1}
  Spectrum 2: –OH stretch at 3000 cm^{-1} and C=O stretch at 1750 cm^{-1}. This is a typical IR spectrum for esters.
  b $CH_3CH_2CH_2COOH$
  c Methyl propanoate
  Ethyl ethanoate
  Propyl methanoate
  d They all have the same number of carbon environments and the chemical shifts would also be similar.
  e Yes. Propyl methanoate is the only one that would produce four peaks and can be identified. The other isomers would produce three peaks in the ^{1}H-NMR spectrum.

29 **a** No. The type of the other atom is not confirmed. Information about proton splitting, chemical shift and integration curves (peak area) would also be useful to determine the number and type of hydrogen atoms.
  b $(CH_3)_3–C–CHO$; there is one more C peak than there are H peaks, meaning that one of the carbon atoms in the chain does not have a hydrogen connected. It is likely that it is connected to three methyl groups and the three methyl groups are equivalent in both C and H environments.

30 $CH_3CH_2COOCH_2CH_3$; ethyl ethanoate

31 $(CH_3)_3CCOOH$; 2,2-dimethyl propanoic acid

32 C_6H_{12} (cyclohexane)

33 **a** C: 5; H: 5
- **b** C: 3; H: 3
- **c** C: 2; H: 2
- **d** C: 5; H: 5
- **e** C: 2; H: 2
- **f** C: 4; H: 3

34 They provide different aspects of information about the molecule, so they can be used in combination to determine the structure of an organic compound. IR provides information on the presence and absence of functional groups. ^{1}H-NMR and ^{13}C-NMR provide information on the number of environments and on how atoms are connected to each other. MS provides information on molecular formula and the potential structure of some groups that can be ionised easily, such as methyl or ethyl groups.

35 **a** Methionine is less polar than serine due to the presence of longer carbon chains. Methionine came out later than serine, so the stationary phase retained non-polar molecules more. The stationary phase is non-polar.
- **b** Changing the concentration of phosphoric acid would change the pH of the environment, which leads to different ionisation states of the molecules. Having different ionisation states changes the polarity of the molecule in the solution, which changes the elution time of molecules.
- **c** Less polar than serine, but more polar than aminobutyric acid, because it elutes after the polar serine but before the non-polar aminobutyric acid with a non-polar stationary phase. The structure can be a type of amino acid with one more carbon than serine and one less carbon in the carbon chain than aminobutyric acid.

36 It is possible with ^{1}H-NMR but not with ^{13}C-NMR, because there is no carbon in the water molecule. The spectrum would show one peak of hydrogen without splitting from 1 to 6 ppm depending on the type of deuterated solvent used.

Chapter 12: Medicinal chemistry

GROUNDWORK

12A A catalyst is a substance that increases the rate of a chemical reaction. Catalysts lower the activation energy for the reaction, but do not participate in the reaction themselves.

12B In an aldehyde, the C of the carbonyl group is a terminal carbon that is also bonded to one C atom and one H atom. In a ketone, the C of the carbonyl group is not a terminal carbon and is bonded to two C atoms.

12C The intermolecular forces are dispersion forces, dipole–dipole attractions and hydrogen bonding. Dispersion forces are the weakest. Dipole–dipole attractions are stronger than dispersion forces. Hydrogen bonding is the strongest type of intermolecular interaction.

12.1 Extracting and purifying plant compounds for medicines

12.1 CHECK YOUR LEARNING

1 Some plants contain bioactive compounds. These are compounds that can affect chemical processes in the human body. The actual compounds that exert these effects are called active ingredients. They must be identified, extracted from the plant and then purified.

2 Caffeine in coffee and catechins in tea can make people more alert. There is a receptor that makes people feel tired and sleepy. The molecules bind to the receptor and prevent the actual metabolite from binding.

3 Secondary metabolites are substances produced by plants to prevent them from being eaten by predators. For example, catechins and caffeine produced by tea plants make them unappetisingly bitter for goats. This prevents goats from eating the plant.

4 Pilocarpine is a bitter-tasing substance in the jaborandi plant that acts as a chemical defence against herbivores. This substance has medicinal uses in humans as it can reduce the pressure inside the eye. It does this by limiting the constriction of the pupil after dilation.

5 **a** For distillation, the boiling point of the substance must be known. It is also very helpful to know the boiling points of the other compounds that you are trying to separate the substance from. For solvent extraction, the substance needs to be soluble in an extraction solvent. Similarly, it would help to understand the solubility of other compounds in the mixture.
- **b** Distillation may be more useful if the substance of interest is a liquid and has a clearly different boiling point compared to other compounds in the mixture. Similarly, solvent extraction may be the better option if the desired substance is soluble in a solvent that will not dissolve the other compounds in the mixture.

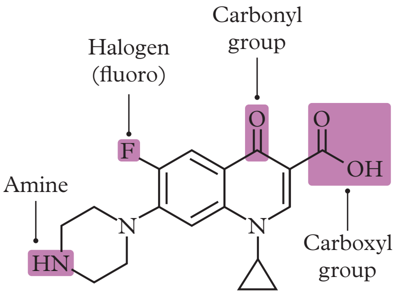
6 **a** The low purity of the patchoulol extract could be due to the fact there are too many compounds in the patchouli plant with a similar boiling point to patchoulol. This means that in the temperature range close to the boiling point of patchoulol, the other compounds are equally volatile and will evaporate, then condense and be collected in the same fraction.
- **b** Student answers will vary but could suggest repeating the distillation process multiple times to improve purity of the patchouli extract.

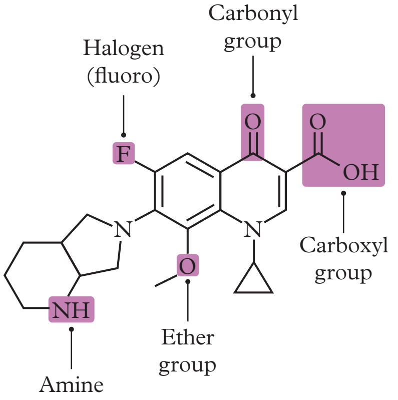
7 Student answers will vary, but should include the steps in the following order: drug discovery, animal studies, clinical trial phase I, clinical trial phase II, clinical trial phase III, drug approval.

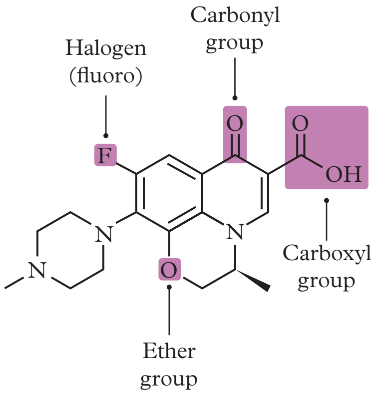
12.2 Identifying the structure and functional groups of medicines

12.2 REAL-WORLD CHEMISTRY

1







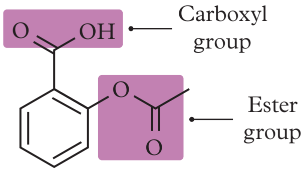
2 Ciprofloxacin and moxifloxacin have the same functional groups in the same locations, except for the amine. By comparing these with levofloxacin, we can see that the three molecules have the halogen, carbonyl group and carboxyl group in common. Therefore, it is likely that these functional groups help the molecule bind to its target in the human body.

3 Student answers may vary but could suggest that they have similar names to make it easy to tell that they are used for the same purpose because they have the same suffix.

12.2 CHECK YOUR LEARNING

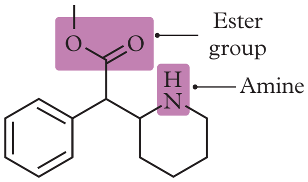
1 The exact configuration of functional groups in an organic compound is important for its medicinal properties because it will allow the compound to bind with its target in the human body.

2



Students may also identify the benzene ring.

3



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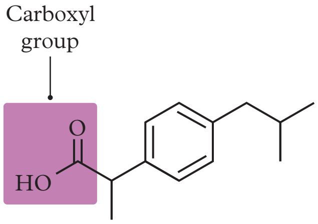
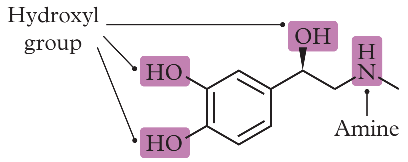
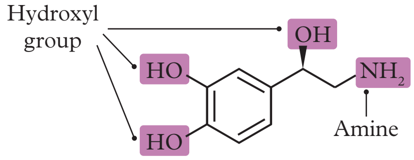


5 a





b Adrenaline and noradrenaline have almost identical structures. It is possible that all functional groups are involved with binding to the target in the human body.

12.3 Chirality in medicine

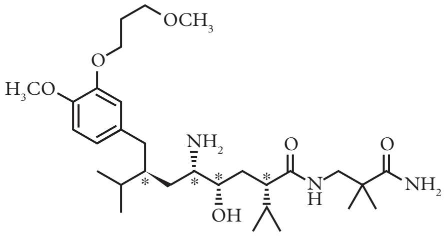
12.3 CHECK YOUR LEARNING

1 A chiral centre is the carbon atom that is connected to four different atoms or groups of atoms.

2 Structural isomers are compounds with the same molecular formula but that contain atoms connected in different ways. This gives them different physical and chemical properties. Stereoisomers have the same molecular formula and contain atoms bonded in the same way. However, they contain one or more chiral centres. This makes their 3D configuration different and can affect how they interact with biological targets.

3 In many cases, only one stereoisomer of a medicinal compound is biologically active, or one stereoisomer works better (more effective, stronger binding to a biological target) than the other(s). Sometimes, the other stereoisomer may be toxic.

4 a



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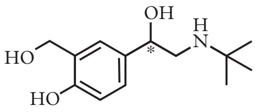
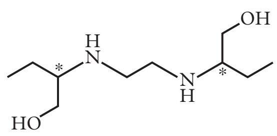


c Fenofibrate does not have any chiral centres.

d



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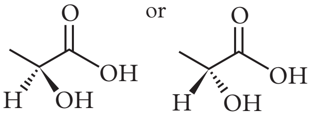
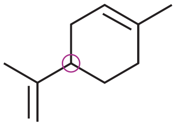


6 a



b Physical and chemical properties (e.g. solubility, melting point, ability to undergo specific chemical reactions)

c 3D configuration and the way they interact with biological targets

12.4 Enzymes as protein-based catalysts

12.4 CHECK YOUR LEARNING

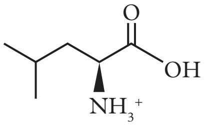
1 Enzymes are protein-based catalysts in living systems that speed up chemical reactions in the body.

2 Refer to your student obook pro.

3 Student answers will vary, but an example is: serine and threonine.

4 Student answers will vary, but an example is: histidine and lysine.

5 a



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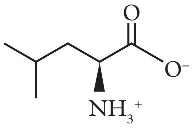

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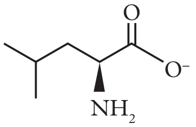


6 Changes in pH can affect the ionisation and therefore charge of some side groups of amino acids. This can affect their three-dimensional structure and interactions with other molecules in the body. Outside of the optimal pH, the enzyme activity is decreased.

7 **a** Decreasing the temperature decreases the rate of any chemical reaction, including chemical reactions catalysed by enzymes.

b Increasing temperature increases the rate of a chemical reaction. However, it also irreversibly changes the three-dimensional shape of enzymes as it destabilises the bonds that comprise the secondary, tertiary, and quaternary structure of the enzyme. The enzyme loses its ability to catalyse chemical reactions.

8 Student answers will vary, but an example response is: To have a quaternary structure, two or more polypeptide chains are required. Some enzymes only require one polypeptide chain to carry out their intended function, so they do not need to form a quaternary structure.

12.5 Medicines that function as competitive enzyme inhibitors

12.5 CHECK YOUR LEARNING

1 Disease often involves problems with chemical reactions in the body that are made possible by enzymes. By blocking the function of these enzymes, we can treat disease. This can be done using competitive enzyme inhibitors, which bind to the active site of enzymes and prevent the natural substrate from binding. Because this might have benefits in treating disease, these inhibitors can be considered medicines.

2 Competitive enzyme inhibitors must have a similar three-dimensional structure to the natural substrate of the enzyme, so that it can bind to the active site of the enzyme. They must also not undergo a reaction after binding.

3 Competitive enzyme inhibitors bind to the active site of an enzyme. This is similar to a lock and key, since the inhibitor must have a complementary three-dimensional shape to the active site of the enzyme.

4 Reversible inhibitors bind to an enzyme's active site with hydrogen bonding, ionic interactions, and dispersion forces, which can be broken relatively easily, releasing the medicine from the active site of the enzyme and allowing the enzyme to catalyse reactions again. Irreversible inhibitors form stronger, usually covalent bonds with the active site of the enzyme, blocking the active site irreversibly.

Chapter 12 review

MULTIPLE CHOICE

1	D	**2**	A	**3**	B	**4**	D	**5**	A
6	C	**7**	D	**8**	B	**9**	C	**10**	B

SHORT ANSWER

11 A molecule containing both a positively and negatively charged group.

12 It provides a bitter taste, which prevents animals from eating it.

13 Boiling point is used for purification by distillation. Solubility is used for purification by solvent extraction.

14 The three-dimensional structure of a medicine allows it to bind specifically to biological targets.

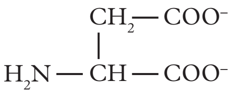
15 One stereoisomer of the medicinal compound may interact with the biological target in a way that exerts a medicinal effect. The other stereoisomers may not be effective or even harmful.

16 **a** Decreasing temperature decreases the rate of any chemical reaction, including chemical reactions catalysed by enzymes.

b Increasing temperature increases the rate of a chemical reaction. However, it also irreversibly changes the three-dimensional shape of enzymes as it destabilises the bonds that comprise the secondary, tertiary, and quaternary structure of the enzyme. The enzyme loses its ability to catalyse chemical reactions because its specific three-dimensional shape has been altered.

c Changes in pH change the charge of some side groups of enzymes. This can affect their tertiary structure and also how they interact with other molecules in the body. This decreases enzyme activity.

17



$$\begin{array}{l} \qquad\;\; \mathrm{CH_2-COO^-} \\ \qquad\;\;\; | \\ \mathrm{H_2N-CH-COO^-} \end{array}$$

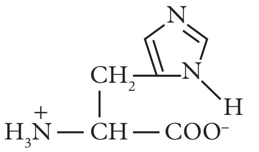
18 Both competitive enzyme inhibitors and the natural substrate for the enzyme must bind to the enzyme by the lock-and-key model. They must have a three-dimensional shape that is complementary to the shape of the enzyme's active site. Because of this, they are often similar in structure.

19 At pH 2: 2+

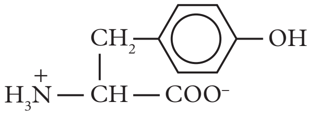
At pH 7: 0

At pH 12: –1

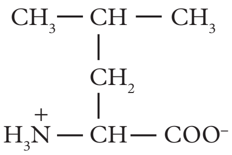
20 **a**



b

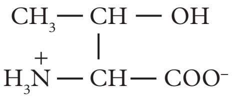


c



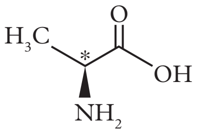
$$\begin{array}{c} \mathrm{CH_3-CH-CH_3} \\ | \\ \mathrm{CH_2} \\ | \\ \mathrm{H_3\overset{+}{N}-CH-COO^-} \end{array}$$

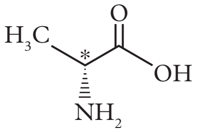
d



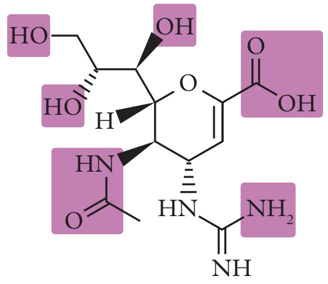
$$\begin{array}{c} \mathrm{CH_3-CH-OH} \\ | \\ \mathrm{H_3\overset{+}{N}-CH-COO^-} \end{array}$$

21

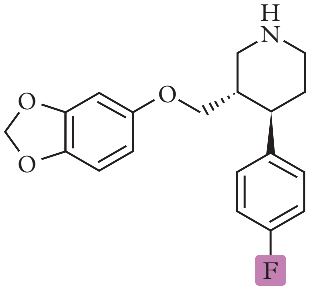




22 **a** I



II



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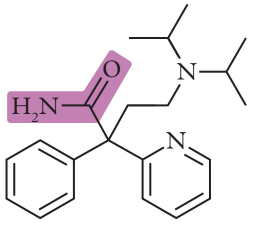
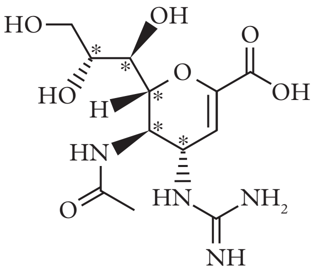
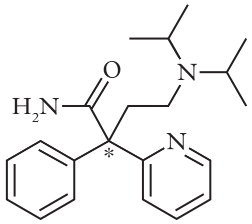


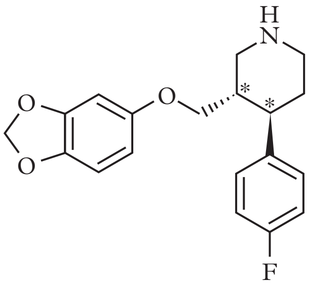
b I



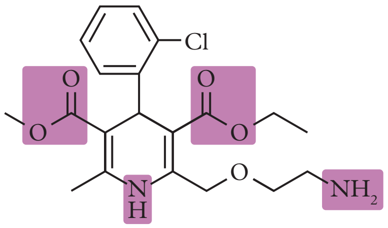
II



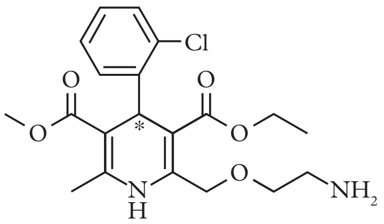
III



23 **a**



b



24 Primary structures are formed from covalent bonds between amino acids. Secondary structures are formed from hydrogen bonding between amide groups on the same polypeptide chain. Tertiary structures are formed from covalent bonds, ionic interactions, hydrogen bonding and dispersion forces between amino acid side groups on the same polypeptide.

25 **a** Reversible inhibitors bind to an enzyme's active site through hydrogen bonding, ionic interactions, and dispersion forces. It competes with the natural substrate for binding at the active site of an enzyme. Less enzyme can bind and the activity is decreased.

b Irreversible inhibitors form stronger, usually covalent bonds with the active site of the enzyme, blocking the active site irreversibly and preventing it from binding with substrate molecules.

26 **a** Student answers will vary, but should suggest dissolving the plant material in ethanol and then using filtration to separate the dissolved compound from the remaining plant material.

b Student answers will vary, but should suggest placing the plant in a solvent, then heating it above the boiling point (230°C) to evaporate the desired compound. The vapour can then be condensed into a liquid and collected.

c The solubility and boiling points of the other compounds in the plant material

27 **a** Student answers will vary but should mention selecting the correct temperature and pH.

b The pH value of the alternative solution is very different from the desired pH. This may affect the structure of the enzyme, since side groups along the amino acid chain may gain or lose a charge and this may affect the interactions between side groups. The second pharmacologist is wrong and conducting the experiment using this other solution is not recommended.

Unit 4 Area of Study 2 Checkpoint

MULTIPLE CHOICE

1	B	**2**	C	**3**	B	**4**	D	**5**	A
6	D	**7**	C	**8**	B	**9**	A	**10**	C

SHORT ANSWER

1 **a** **i** Molecule F

ii Looking at the number of peaks, it could be molecule F or G. However, the last peak at >200 ppm must correspond to a C=O bond in a ketone. The peak for an ester carbon will appear lower on the spectrum at 165–175 ppm. Therefore, it cannot be molecule G.

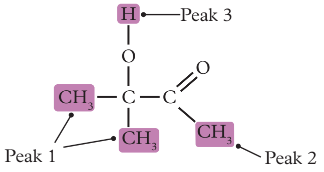
b **i** Molecule G

ii Spectrum A is produced by an ester. Since molecules D, E and F contain –COOH and –OH groups, they would produce large absorption bands in the 2500–3500 cm^{-1} range, which are not observed in Spectrum A.

iii You cannot definitively determine whether Spectrum B belongs to molecules D or E. Both molecules have an –OH functional group and will therefore produce the same absorption band at 3200–3600 cm^{-1}. They cannot be differentiated from the IR spectrum alone.

c **i** Molecule E

ii



iii Molecule D: 2 peaks; molecule F: 5 peaks; molecule G: 4 peaks

d You could use mass spectrometry to determine the difference. The molecular ion peak will be the same for all molecules. You would have to compare the base peak and other significant peaks in the spectrum to determine the difference between the molecules.

2 **a** We could expect molecule A to have the lowest molar mass, followed by B and then C, with the highest molar mass. However, the order of elution will also depend on molecule polarity. The more polar molecule will have greater affinity for the stationary phase and a longer retention time.

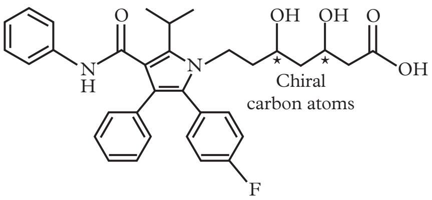
b Molecule C is likely to have the greatest polarity and has greater affinity for the stationary phase.

c The athlete's blood must contain molecule B.

d 350 mg L^{-1}.

3 **a** Alkene, hydroxyl group, carboxyl group, fluorine (or halogen)

b **i**



ii Stereoisomers contain the same number and type of atoms, connected in the same way, but arranged in a different three-dimensional configuration. There are four stereoisomers of atorvastatin, shown in your obook pro.

c Often, only one stereoisomer can bind well with the target receptor or enzyme because of the three-dimensional shape of the stereoisomer and the complementary shape of the receptor or enzyme.

4 **a** Tertiary structure

b A – dispersion forces; B – hydrogen bonding; C – covalent disulfide bridges; D – dispersion forces

c 38°C (acceptable range 36–39°C)

d It will function much slower than at its optimal temperature but will not denature.

e They can have a similar structure to that of the target enzyme substrate. They will bind to the active sites of the enzymes instead of the target substrate.

f Irreversible competitive inhibitors can form strong bonds with the active site to block the active site irreversibly. On the other hand, a reversible competitive inhibitor interacts with the active site through hydrogen bonds, ionic interactions or dispersion forces, which are not permanent.

Unit 4 review

MULTIPLE CHOICE

1	C	**2**	D	**3**	D	**4**	B	**5**	B
6	A	**7**	B	**8**	D	**9**	D	**10**	A

SHORT ANSWER

1 **a** Alkanes are held together by weak dispersion forces, which require less energy to break compared with molecules that can form strong hydrogen bonds, such as alcohols and carboxylic acids.

b With an increase in the number of carbon atoms, there are more electrons in the molecule and therefore more opportunities for temporary dipoles to form, causing the dispersion forces between molecules to increase. The greater the dispersion forces, the higher the boiling point.

2 The full pathway can be found in your obook pro.

3 **a** **i** 1

ii 2

iii 1

b **i** Stereoisomer

ii Same

4 **a** $C_3H_6O_2$

b 29: CH_2CH_{3+}

45: $COOH_+$

73: $CH_3CH_2COO_+$

74: $CH_3CH_2COOH_+$

c CH_3CH_2COOH

d **i** O–H (acid)

ii C–H

iii C=O

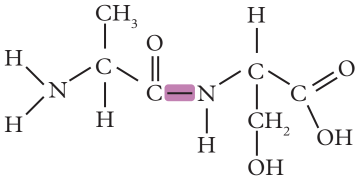
e Propanoic acid

5 **a** A high temperature needs to be used. A sufficient amount of ethanol is produced at 300°C, but the assistance of a catalyst is required to speed up the reaction.

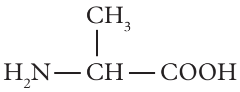
b The forward reaction would be favoured, as the equilibrium will try to re-establish itself by favouring the production of ethanol.

c Energy is required to obtain high pressures, which increases the costs.

6 **a**

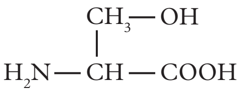


b Alanine



$$\begin{array}{c} CH_3 \\ | \\ H_2N-CH-COOH \end{array}$$

Serine



$$\begin{array}{c} \quad\; CH_3-OH \\ | \\ H_2N-CH-COOH \end{array}$$

7 Because water is a polar molecule, only polar substances can dissolve in water. Hexane and hexene are non-polar hydrocarbons, so they cannot dissolve in water. Ethanol and propanoic acid are polar molecules that can form hydrogen bonds with water molecules and hence can dissolve in water.

GLOSSARY

^{1}H-NMR
hydrogen-1 (or proton) nuclear magnetic resonance spectroscopy; the analysis of hydrogen-1 nuclear spin to identify hydrogen environments in a compound

^{13}C-NMR
carbon-13 nuclear magnetic resonance spectroscopy; the analysis of carbon-13 nuclear spin to identify carbon environments in a compound

A

Aboriginal and/or Torres Strait Islander Peoples
people with family heritage from, and membership in, the ethnic groups that lived in Australia before British colonisation

absorption bands
peaks in an IR spectrum

abundance
how often the substance occurs

acronym
an abbreviation formed from the initial letters of each word from a group of words

accuracy
how close the value is to the true value

activation energy
the minimum energy required to break the bonds between atoms in reactants

active ingredient
a compound that is bioactive; in medicine, it is the compound that has medicinal properties

active site
the region of an enzyme that binds one or more substrates

actual yield
the final measured mass of product after completing a reaction

addition reaction
a reaction in which two molecules add together to form one new molecule; usually involving a double bond

adenosine triphosphate (ATP)
the organic molecule that provides energy for metabolic processes in an organism

adsorption
the process in which a component becomes attracted to the stationary phase

aeration
the process in which a substance is exposed to air

aerobic respiration
cellular respiration in the presence of oxygen

aim
a statement that describes the purpose of an investigation

aldehyde
a molecule that contains a carbonyl group on a terminal carbon atom

aliquot
the volume of solution dispensed, using a pipette, into the conical flask

alkane
a saturated hydrocarbon that only contains single covalent bonds

alkyl chain
the general name for a main hydrocarbon chain consisting of any number of carbon atoms

alkyl group
the general name for a hydrocarbon branch consisting of any number of carbon atoms

alpha helix
a coiled structure formed due to hydrogen bonding between the amide groups on a single polypeptide chain

amine
a homologous series that contains the functional group NH_2

amino acid
a small organic compound which contains a $-NH_2$ and $-COOH$ functional group

amino functional group
an NH_2 molecule that will attach to a carbon atom to form an amine

amylopectin
a type of plant starch that is branched in structure

amylose
a type of plant starch that is linear in structure

anaerobic fermentation
cellular respiration in the absence of oxygen

anode
the negatively charged electrode

aqueous
the state of a substance that has been dissolved in water

aqueous reaction
the reaction of two solutions that form aqueous products

aromatic hydrocarbon
a hydrocarbon compound that contains a ring with delocalised electrons

artificial photosynthesis
the production of hydrogen gas by water splitting and proton catalyst reduction using light energy

atom economy
the ratio of the amount of product yielded to the amount of reactants used in a reaction; a measure of the efficiency of a reaction in converting reactants to products

B

B20
a fuel blend of less than 20% biodiesel and more than 80% petrodiesel

base peak
the peak on the mass spectrum that has the highest relative intensity; it is the most abundant and stable ion present

Benedict's solution
a mixture of copper(II) sulfate pentahydrate, sodium carbonate and sodium citrate

beta-pleated sheet
a folded structure formed due to hydrogen bonding between the amide groups on a single polypeptide chain

bioactive
can bind to targets and affect chemical processes in the human body

bio-based polymers
biomass-derived polymers or plastics

biodegradable plastics
polymers or plastics that can be broken down by living organisms

biodiesel
a renewable fuel consisting of a mixture of fatty acid methyl esters produced by the transesterification reaction between plant oils and methanol or ethanol

bioethanol
ethanol produced by the fermentation of sugars (e.g. glucose) by yeast enzymes, then distilled to a high enough purity (>95%) to be used in an internal combustion engine

biofuel
fuels produced from plants

biogas
a fuel consisting of a blend of gases including methane and carbon dioxide, produced by the anaerobic fermentation of plant waste in a biogas reactor vessel

biological target
a target (e.g. a receptor or enzyme) in a living system (e.g. the human body) to which a compound can bind and exert a biological effect

biomolecule
a molecule produced by living cells and organisms

boiling point
the temperature at which a liquid becomes a gas

bond energy
the amount of energy (in kJ mol^{-1}) that it takes to break a covalent bond

bond stability
the more energy that is required to break the bond, the more stable the bond is

bond strength
how strongly two atoms are joined together

bromine test
a qualitative test in which bromine is added to a hydrocarbon to test for carbon–carbon double bonds; if the bromine changes from red–orange to colourless, the hydrocarbon is unsaturated

burette
a graduated glass cylinder that dispenses high-accuracy volumes of standard solution

by-product
a secondary product formed in a chemical reaction

C

calibration curve
a graph of known concentrations and peak areas from HPLC; used to identify an unknown concentration

calibration curve (in calorimetry)
a graph of temperature versus time used to determine the efficiency of a calorimeter

calibration factor
the amount of energy that it takes to increase the temperature of an aqueous substance in a calorimeter by 1°C; this value is used to calibrate the calorimeter

calorimeter
a device that measures the change in temperature of water when an endothermic or exothermic reaction takes place

carbohydrate
a macronutrient consisting of one or more sugar units

carbon environment
the atoms and bonds that surround a specific carbon atom

carbonyl group
a carbon atom double-bonded to an oxygen atom

carboxyl group
a hydroxyl group bonded to a carbon atom that has a double-bonded oxygen atom

carboxylic acid
contains a carboxyl group

catalysis
the acceleration of a chemical reaction by a catalyst

catalyst
a substance that lowers the activation energy of a reaction and increases the rate of reaction without participating in the chemical reaction

cathode
the positively charged electrode

cell diagram
a diagram of a galvanic cell that has all the key components, reaction mechanisms and electron/ion movements labelled

cell potential difference
the difference in electrical potential between two electrodes, represented by E_0 cell; the voltage (in volts, V) that the galvanic cell can generate at standard conditions

cellular respiration
a combustion reaction in which nutrient molecules react with oxygen to produce energy

cellulose
a long-chain linear polysaccharide made up of β-glucose monomers; not digestible by humans

chain isomer
a structural isomer caused by branching of the hydrocarbon chain

chemical additive
a chemical that can be introduced into a system to improve a reaction

chemical energy
energy stored within chemical bonds

chemical environment
the atoms and bonds that surround a specific atom

chemical purity
a measurement of the absence of impurities in a sample

chemical shift
the amount of energy required to spin-flip a nucleus compared with TMS

chemistry
the study of matter, including its structure, properties and behaviour

chiral centre
a carbon atom with single covalent bonds to four different atoms or groups of atoms

chloroplast
the structure in living cells in which photosynthesis occurs

circular economy
a model of production that involves sharing, leasing, re-using, repairing, refurbishing and recycling existing materials and products as long as possible

climate change
the gradual increase in the temperature of the Earth's surface, oceans and atmosphere; generally caused by the greenhouse gas effect

closed system
a reaction vessel that is closed off, or contains a lid, meaning that no reactant or product particles can escape

collate
collect and combine texts, information or data

collision theory
the theory that states that reactants must collide in the correct orientation and with sufficient energy for a reaction to occur

combustion
a reaction with oxygen that produces carbon dioxide, water and heat energy

competitive enzyme inhibitor
a compound that binds to the active site of an enzyme and prevents the binding of the natural substrate

complete combustion
the burning of a fuel in excess oxygen supply to produce carbon dioxide gas and water vapour

complex half-equation
a half-equation for an oxidation and reduction equation that occurs under acidic or basic conditions

concentration
a measure of the number of particles in a volume of substance

concordant titres
titres within 0.1 mL from the highest to lowest volume

condensation
the transition of a substance from gaseous to liquid state

condensation reaction
a reaction in which two molecules are added together to form a new larger molecule and a water molecule, or another small molecule

conjugate redox pair
two chemical species that differ by the number of electrons that are transferred in the redox reaction

continuous data
data or information that can be any numerical value

controlled variable
a variable that is not changed or is constant during an investigation

correlation
the presence of a relationship between two variables; can differ in strength

counterfeited
an inauthentic imitation of a substance

counting prefixes
used to indicate multiple identical branches or functional groups; 2 = di-, 3 = tri-, 4 = tetra-

covalent bond
a bond formed when atoms share electrons; can only occur between two non-metals

current
the flow of electrons generated by a power source; measured in amps (A) and represented by I

cyclic
a compound that contains atoms connected in a ring

cycloalkane
a hydrocarbon compound containing only single C–C bonds, in which the carbon chain closes to form a ring

D

data evaluation
the critical analysis of data that has been personally collected or provided to identify contradictory or incomplete data, or issues such as a personal bias

degree of unsaturation
the number of double bonds or rings in the structure

delocalised
(when referring to electrons) not associated with a single atom or bond and therefore free to move around the chemical structure

denaturation
the process in which an enzyme loses its three-dimensional structure and can no longer bind to its natural substrates

dependent variable
the variable that is observed or measured when the independent variable is changed during an investigation

deposition
the transition of a substance directly from gaseous to solid state

desorption
the process in which a component is released from the stationary phase back into the mobile phase

dimer
two monomers joined by intermolecular bonds

dipole
unequal sharing of electrons creating a slight negative or positive charge at each end of the bond

discharge
the conversion of chemical to electrical energy

discrete data
data or information that can only be certain numerical values

dissolution
the process of dissolving a solute into a solvent (typically water) to create a solution

distillation (of a plant)
a process by which a desired compound is separated from a mixture by using the fact it has a different boiling point to the other components in the mixture

distillation (purity)
a separation technique where a mixture is vaporised, rises and passes through a condenser where each component is turned back into a liquid at a different point, and then collected, having been separated from the mixture

dynamic equilibrium
the state a reaction reaches when the forward and reverse reactions are occurring at the same rate

E

E10
a fuel blend of up to 10% bioethanol and 90% or more of petrol

electric charge
the rate of flow of electrons generated by a power source; measured in coulombs (C) and represented by Q

electrochemical cell
a device in which electrical energy can be generated from chemical energy, or chemical energy from electrical energy

electrochemical series
a table of oxidising agents and reducing agents written in reduction half-equations, where the strongest oxidising agent is at the top left of the table and the strongest reducing agent is at the bottom right

electrode
a strip or rod that conducts electricity; this is the surface on which the oxidation or reduction reactions occur

electrode polarity
the charge of an electrode

electrolysis
a process in which electrical energy is passed into a cell, using a power source, to reverse a spontaneous redox reaction

electrolyte solution
an aqueous solution containing ions

electrolytic cell
a type of electrochemical cell where electrical energy is converted to chemical energy

electromagnetic radiation
the entire wavelength region on the electromagnetic spectrum, including UV and visible light, infrared radiation, microwaves and radio waves

electromagnetic spectrum
the range of wavelengths that all electromagnetic radiation exists at, including UV and visible light, infrared radiation, microwaves and radio waves

electron
a subatomic particle with a negative charge

electronegativity
the ability of an atom in a bond to attract a bonding pair of electrons

electroplating
the process in which one metal is coated with another using an electrolysis reaction

electroplating cell
an electrolytic cell that converts electrical energy to chemical energy in order to coat an object in a layer of metal

electropurification
the process in which a pure metal is extracted from an impure mixture using an electrolysis reaction

electrosynthesis
a process in which electrolysis is used to form new products

eluate
the liquid that has left the chromatography column; may contain one or more sample components dissolved in the mobile phase

elute
exit of the chromatography column

end point
the change in colour of the indicator during a titration

endothermic
describes a chemical reaction in which heat energy is absorbed from the surroundings, and the enthalpy of the reactants is lower than the enthalpy of the products

energy content
the total energy released that can be used by the body when nutrient molecules are combusted

energy profile diagram
a graph that shows the change in enthalpy as a reaction progresses

enthalpy
a measure of chemical energy; represented by H

enthalpy of combustion
the change in enthalpy when a quantity of fuel undergoes complete combustion

enzyme
a biological, protein-based catalyst that speeds up chemical reactions

enzyme–substrate complex
the structure formed when one or more substrates bind to the active site of an enzyme

equilibrium arrows
reaction arrows that point in both directions to indicate that a reaction is reversible

equilibrium constant (K)
the ratio of reactants to products in a reaction when it is at equilibrium

equilibrium expression
the non-numerical representation of the equilibrium law, which states the chemicals to the power of their coefficients

equilibrium law
the concentration of products to the power of their coefficients, divided by the concentration of products to the power of their coefficients, is equal to the equilibrium constant, K

equilibrium reaction
a reaction that is reversible; reactants and products are both present in the system at any one time

equivalence point
the point in a titration when the chemicals have reacted in the molar ratio of the balanced chemical equation

error
the difference between an accepted or theoretical value and the experimental, observed or measured value

ester
a compound produced as the result of a reaction between an alcohol and a carboxylic acid

esterification
a condensation reaction to form an ester and a water molecule from an alcohol and carboxylic acid

ether link
–O– bond that links together two carbon atoms

ethics
moral principles that govern a person's behaviour or how an activity is conducted

excess reagent
leftover reactant in a reaction; not used to calculate the final amount in stoichiometry

exothermic
describes a chemical reaction that is accompanied by the emission of heat energy to the surroundings; if the enthalpy of the

products is lower than the enthalpy of the reactants, the reaction loses energy as heat to the environment

experimental yield
the actual amount of product formed in a chemical reaction; also called actual yield

extent of a reaction
a measure of the magnitude to which a reaction proceeds in the forward direction

external circuit
the electric circuit (wires) which allow for electron movement from anode to cathode; also includes the voltmeter

F

Faraday constant
1 F = 96 500 C mol^{-1}; 96 500 coulombs are required to generate 1 mol of a metal or gas at the cathode of an electroplating cell

Faraday's first law
the mass of a metal plated onto a cathode is proportional to the charge applied to an electroplating cell

Faraday's second law
to produce 1 mol of a metal, 1, 2, 3 or another whole number of moles of electrons must be consumed

feedstock
a raw material used to produce other substances

fingerprint region
the absorption pattern in the wavenumber range 500–1400 cm^{-1}; unique to each compound

forward reaction
the reaction between reactants to form products

fossil fuel
a non-renewable fuel, such as coal, crude oil and natural gas, obtained from decomposed organic matter that has been compressed under heat and pressure for millions of years underground

fraction
in distillation, a substance or mixture of substances with the same or a similar boiling point

fractional distillation
a separation technique where components of a mixture are vaporised, rise through a column until they reach a compartment of a temperature lower than their boiling point, condense and are separated and removed from the column

fragmented ions
the smaller pieces formed when the bonds holding a molecular ion together are broken

freezing
the transition of a substance from liquid to solid state

frequency
the number of times an observation occurred in an experiment or a study

frequency of collisions
how often collisions occur between reactants (including successful and unsuccessful collisions)

fuel
a substance that reacts to produce heat energy

fuel cell
a galvanic cell that produces electricity by using a constant supply of reactants (often hydrogen and oxygen) and inert electrodes that do not break down

functional group
a specific group of distinctly arranged atoms that changes the properties of a molecule

functional group region
the absorption pattern in the wavenumber range above 1400 cm^{-1}; provides information about the functional groups present in the compound

G

galvanic cell
a type of electrochemical cell where the redox reaction is separated into reduction and oxidation half-cells; these are connected using a circuit to generate electricity

glycogen
a highly branched polysaccharide made up of glucose monomers; used for energy storage in animals

glycosidic bond
an ether link between two saccharides

green chemistry
an area of chemistry that focuses on designing safer and more sustainable new products

'green' hydrogen
hydrogen produced by chemical techniques that are efficient, require minimal energy, produce no waste and are carbon-neutral

greenhouse effect
a phenomenon whereby greenhouse gases (e.g. methane, carbon dioxide, water vapour) trap heat in the atmosphere and cause Earth to warm

greenhouse gases
gases such as CO_2, H_2O and CH_4, which contribute to the greenhouse effect

group
a vertical column on the periodic table

H

half-cell
the component of a galvanic cell that contains either the oxidation or reduction redox reaction

half-equation
an equation which represents half of a redox equation and is either an oxidation or reduction half-equation; it shows electron transfer

heat of combustion
the amount of heat energy released when a specific quantity of a fuel combusts

heterogeneous catalyst
a catalyst that is in a different state from the reactants

heterogeneous equilibrium
an equilibrium reaction where the states of reactants and products are not the same

high performance liquid chromatography (HPLC)
the technique used to separate, identify and calculate the amount of each component in a compound

high resolution
(in NMR spectroscopy) spectra that show finer details for each peak, such as peak splitting

homogeneous catalyst
a catalyst that is in the same state as the reactants

homogeneous equilibrium
an equilibrium reaction where the states of all reactants and products are the same

homologous series
a sequence of compounds that share the same characteristics; also known as hydrocarbon families

hydrocarbon
a compound of carbon and hydrogen atoms

hydrogen environment
the atoms and bonds that surround a specific hydrogen atom

hydrogen–halide
a molecule containing a hydrogen and a halogen; can be H–Br, H–Cl, H–F or H–I

hydrogenation
the addition of H_2 to an alkene to form an alkane

hydrolysis reaction
the reverse of a condensation reaction – water is added to a reactant to form two or more new molecules

hydrometer
an instrument used to measure the density of a substance

hydrophobic
has a low affinity for water; generally non-polar

hydroxyl group
an oxygen and a hydrogen atom (–OH) attached to an alkyl chain or group R

hypothesis
a testable statement that includes a prediction about the outcome of an investigation based on scientific reasoning

I

incomplete combustion
the burning of a fuel in limited oxygen supply to produce carbon monoxide gas, solid carbon (soot) and water vapour

independent variable
the variable that is changed or manipulated during an investigation

indicator
a chemical substance that can be chosen to change colour when the equivalence point of a reaction is reached

inert electrode
an electrode in a half-cell that does not participate in the reaction – i.e. it is neither a reactant nor a product

infrared radiation
a wavelength region on the electromagnetic spectrum of 700–1000 nm

infrared spectroscopy
an analytical technique in which bond type can be identified by measuring the amount of infrared radiation absorbed by the bonds

instrument error
an error due to the limited accuracy of a scientific instrument; usually very small (and not significant) if used with standard operating and/or laboratory procedures

integration curve
a curve in which the height corresponds to the relative peak area

intercalated
inserted between layers

intermolecular forces
the attractive forces between covalently bonded molecules, including hydrogen bonding, dipole–dipole attractions and dispersion forces

internal circuit
the salt bridge that allows ions to flow

iodine test
a qualitative test in which iodine is added to a hydrocarbon sample to test for carbon–carbon double bonds; if the iodine changes from brown to colourless, then the hydrocarbon is unsaturated

iodohydrocarbon
an organic compound consisting of carbon, hydrogen and iodine atoms

irreversible inhibitor
an inhibitor that binds permanently to an enzyme

irreversible reaction
a reaction that can only proceed in the forward direction

isolated system
a thermodynamic system that cannot exchange matter or energy outside its boundaries

isotopes
different forms of the same element with the same number of protons but different numbers of neutrons

J

jargon
special words or expressions used by a profession or group that are difficult for others to understand

K

ketone
a molecule that contains a carbonyl group on a carbon atom along the main chain

kinetic molecular theory
the theory that states that all particles are in constant random motion

L

law of conservation of mass
a law that states that in an isolated system, mass cannot be created or destroyed

Le Chatelier's principle
if the position of equilibrium in a system is subjected to change, the system will shift to partially oppose the effect of the change

like dissolves like
a rule stating that polar molecules will dissolve better in polar solvents and non-polar molecules will dissolve better in non-polar solvents

limiting reagent
the reactant that is completely used up in a reaction; used to calculate the final product amount in stoichiometry

line of best fit
a line drawn through a scatterplot of data points that expresses a relationship between those points

linear economy
a way of managing resources that operates on a take–make–dispose model; new resources are used and disposed of after use

lipid
a macronutrient that has a very low solubility in water

low resolution
(in NMR spectroscopy) spectra that show only single peaks

M

macronutrient
a nutrient needed by the body in large amounts; macromolecules that provide energy to the body

main chain
also known as the parent chain, parent molecule or longest chain

mass spectrometry
an analytical technique in which the masses and relative abundances of isotopes, molecules and molecular fragments are measured by ionising them and determining their trajectories in electric and magnetic fields

mass spectrum
a column graph that shows the relative abundance of each ionised fragment in a sample

mass-to-charge ratio (*m/z*)
the molar mass of the positive ion formed divided by its charge

measurement result
the final result reported in an investigation

melting
the transition of a substance from solid to liquid state

melting point
the temperature at which a solid becomes a liquid

methodology
a system of methods used in a particular area of study or activity

mobile phase
the liquid that a chromatography sample is dissolved into, which then moves through the HPLC

molar heat of dissolution
the enthalpy of a solution when a solid substance dissolves in water

molar heat of neutralisation
the enthalpy of a solution when an acid reacts with a base

molar heat of solution
the amount of energy absorbed or released by a solution

molecular ion
the unfragmented, ionised (positively charged) form of the original molecule

molecular ion peak
the peak on the mass spectrum that corresponds to the positively charged, unfragmented, whole, original molecule

molten
the state of a substance that has been melted to form a liquid

monosaccharides
molecules made up of single saccharide units

multiplet
a hydrogen NMR peak that has split into five or more fine peaks

N

$n + 1$ rule
a rule stating that hydrogen NMR peaks will split into finer peaks; the number of finer peaks = the number of neighbouring hydrogen atoms + 1

net forward reaction
a shift in the position of equilibrium in which more products are being formed compared to reactants being re-formed

net mass
the specific mass of the desired substance(s), excluding its container

net reverse reaction
a shift in the position of equilibrium in which more reactants are being re-formed compared to products being formed

neutron
a subatomic particle in an atom's nucleus with a neutral (or zero) charge

non-cyclic
describes a compound that contains atoms connected only in single or branched chains

non-renewable
takes a very long time to form; is used more quickly than it can be produced

nuclear charge
the total charge of the nucleus due to the number of positively charged protons

nuclear magnetic resonance (NMR)
an analytical technique in which the chemical environments in a compound can be determined by measuring the energy released from the nucleus of an atom after it absorbs electromagnetic radiation

nuclear spin
the overall rotation of an atom's nucleus

nucleic acid
a biomolecule that makes up the genetic material

number of successful collisions
the total number of reactant particles in the reaction that are colliding successfully and forming products

O

open system
a reaction vessel that has no lid, meaning that reactants or products can be lost to the atmosphere

organic compound
a chemical compound that contains carbon–hydrogen bonds

overall equation
an overall redox reaction that combines the two half-equations after electrons have been balanced and cancelled out

OX Z RED
an acronym to remember that oxidation is the top half-equation on the electrochemical series, the Z dictates the reactants and products, and reduction is the bottom half-equation

oxidation
a process in which electrons are lost

oxidation number
the degree of oxidation of a chemical species that indicates the number of electrons gained or lost by the species to bond in a compound; also known as oxidation state

oxidising agent
a reactant species that causes another to lose electrons and be oxidised; is itself reduced

P

parallax error
an error in taking a value made by not having your eye directly in line with the measurement

peak area
the area under a peak

peak splitting
the further splitting of individual peaks caused by neighbouring hydrogen atoms

PEM electrolysis
water electrolysis with an acid electrolyte, used to produce hydrogen gas using a polymer electrolyte membrane

peptide bond
a $CONH_2$ bond; also known as an amide bond

percentage error
the calculated percentage difference between an observed or measured value and an accepted or theoretical value

percentage purity
the percentage of the desired product contained in a product

theoretical value
the values that have been determined using specialised equipment and chemicals of the highest purity; also known as a literature value

percentage yield
comparison of actual yield and theoretical yield from a completed reaction

period
a horizontal row on the periodic table

periodic table
the table that organises the elements in order of increasing atomic number

petrodiesel
diesel fuel obtained from crude oil; a fossil fuel

photoelectrochemical cells
electrochemical cells that use light as a reactant to produce hydrogen gas

photon
an uncharged particle that is the smallest possible amount of electromagnetic or light energy

photosynthesis
the process in which light energy is converted to chemical energy

pipette
a glass tube that dispenses high-accuracy volumes of a solution with unknown concentration

polysaccharides
long-chain molecules made up of multiple saccharides

porous membrane
a barrier that contains holes (pores) for substances to pass through

position of equilibrium
the relative amounts of reactants and products in a reaction system at equilibrium

positional isomer
a structural isomer where the functional group is in different positions on the same carbon chain

precision
how close the data values in a set are to each other

prefix
a letter or group of letters added to the beginning of a word to change its meaning or create a new word

pressure
the force exerted, per unit area, by one substance upon another substance

primary amide
a substance that contains a carbon atom with a double-bonded oxygen atom and NH_2 group attached to the alkyl chain or group R

primary amine
an amine where nitrogen is always bonded to one carbon and two hydrogen atoms

primary cell
a non-rechargeable battery; a galvanic cell that produces an electrical current

primary standard
the solute which is dissolved in a solvent to create a standard solution

primary structure
the structure formed when amino acids are joined in a sequence by covalent amide bonds or peptide linkages

procrastination
the action of delaying or postponing something

proportion of successful collisions
the proportion or percentage of reactant particles in a reaction that are colliding successfully and forming products

protein
a macronutrient consisting of a chain of amino acids

proton
a subatomic particle with a positive charge, found in an atom's nucleus

Q

qualitative data
data that is not numerical, can be text, images or audio

qualitative testing
tests that gather data that is not measurable; often based on observations or descriptions

quantitative data
data that can be counted or measured and is expressed as numbers

quantitative test
a test that gathers measurable data, expressed numerically

quaternary structure
the structure formed when covalent bonds, hydrogen bonds, ionic interactions or dispersion forces form between side groups of amino acids of two or more polypeptide chains

R

racemic mixture
a mixture that contains an equal concentration of stereoisomer pairs

radio wave
a wavelength region of the electromagnetic spectrum above 10 cm

reactant
a chemical species that undergoes change and is consumed during a reaction

reaction quotient (Q)
the ratio of reactants to products in a reaction when it is not at equilibrium

reaction rate
the change in concentration of a reactant or product per unit of time (change over time)

reactive electrode
in a half-cell, an electrode that participates in the reaction and is oxidised during the discharge reaction

reading error
an error made by approximating a measurement if it falls between the markings of a measuring device

reagent
a chemical species that participates in a reaction, but is not necessarily consumed

recharge
conversion of electrical to chemical energy in a redox reaction

RED SOX
an acronym used to remember that reduction is the top half-equation on the electrochemical series; the S dictates the reactants and products, and the bottom half-equation is oxidation

redox
a chemical reaction involving the transfer of electrons from one reactant to another

redox titration
a titration that involves redox reactions

reducing agent
a reactant species that causes another species to gain electrons and be reduced; is itself oxidised

reduction
a process in which electrons are gained

relative intensity/abundance
the percentage of ions produced by the mass spectrometer

renewable
replenishable by natural processes in a relatively short amount of time

repeatability
whether the same data values can be produced again by the same experiment, under the same laboratory conditions

reproducibility
whether the same data values can be reproduced under slightly different conditions

resolution
the smallest increment a tool can measure

retention time (R_t)
the time that a component spends inside a chromatography column before it elutes

reverse reaction
the backward reaction where the products react to re-form the reactants

reversible inhibitor
an inhibitor that does not permanently bind to the enzyme

reversible reaction
a reaction that can proceed in both forward and reverse directions

risk assessment
a process of evaluating the potential risks that may be involved in an activity, e.g. performing an experiment

S

saccharides
sugars

salt bridge
internal circuit connecting two half-cells; allows for the flow of charge by moving ions; cations move to the cathode and anions move to the anode

saturated (molecule)
a molecule that contains only single carbon bonds (C–C)

scientific method
a series of steps used to acquire knowledge in science, which involves observation; developing and testing hypotheses; and collecting, analysing and communicating results

secondary amide
an amide in which the nitrogen atom is attached to one hydrogen atom and one R group

secondary cell
an electrochemical cell that can operate in both discharge (as a galvanic cell) and recharge (as an electrolytic cell)

secondary metabolite
a chemical substance produced by a plant that protects it from predators (e.g. herbivores); typically does not directly affect the growth of the plant

secondary structure
the structure formed when amide groups along a polypeptide chain form hydrogen bonds with each other; forms an alpha helix or beta-pleated sheet

self-indicating
changes colour at the equivalence point, without the need of an additional indicator

semipermeable
describes a membrane that allows some, but not all, substances to pass through

sensitive
(in analytical chemistry) able to detect very low quantities of compound

shielding effect
the shielding of valence electrons by inner electrons that alters the nuclear charge felt by valence electrons

significant figures
the digits of a number that are used to express it to the required degree of accuracy, starting from the first non-zero digit

simple distillation
a separation technique where components of a mixture are vaporised, condensed as they are cooled by running water and collected in a distilling flask

SLC
standard laboratory conditions of 25°C and 100 kPa

solution calorimetry
an analytical technique in which the energy generated by a sample is determined by measuring the change in the temperature of water

solvent extraction
a process by which a desired compound is separated from a mixture because it is soluble in a solvent, but the other components in the mixture are not soluble

specific heat capacity
a measure of the amount of heat energy in joules (J) or kilojoules (kJ) that it takes to increase the temperature of 1 g of a substance by 1°C

spectator ion
an ion that exists in the same form on the reactant side and product side of a redox reaction; there is no change in its oxidation number or state

spirit burner
a vessel with a wick that can hold liquid fuels and combust them in the presence of oxygen in the air

standard cell potential
the same as cell potential difference

standard conditions
the conditions at which all E_0 values have been measured; 25°C, 1 atm and 1 M

standard electrode potential
the electrical potential generated at the electrode under standard conditions; measured in volts (V)

standard solution
a solution with a highly accurately known concentration

starch
the stored form of glucose in plants

stationary phase
the solid within the column of the HPLC onto which the components absorb

steam methane reforming
a process in which methane is heated with steam and a catalyst to produce carbon monoxide and hydrogen gas; also called steam methane cracking (SMC)

stereoisomers
molecules with the same molecular formula that contain atoms connected in the same way, but with a different three-dimensional configuration

stoichiometric ratio
the molar ratio in which the reactants react with each other (the coefficients in the balanced chemical equation)

stoichiometry
calculations using mole ratios for a chemical reaction

structural isomer
a molecule with the same molecular formula as another molecule, but different structure

sublimation
the transition of a substance directly from solid to gaseous state

substitution reaction
a reaction where an atom in one molecule is replaced by an atom from another molecule; if the reaction starts with two reactants, it will end with two new products

substrate
a reactant in an enzyme-catalysed reaction

successful collision
when reactants collide with sufficient energy and in the correct orientation to form products

suffix
a letter or group of letters added to the end of a word to change its meaning or create a new word

surface area
a measure of the total surface area available to react in a chemical reaction

surroundings
the environment in which a chemical reaction is occurring

sustainability
using natural resources more efficiently, to ensure supply of resources continues into the future

system
the chemical bonds holding together each reactant and product molecule

T

terminal carbon
a carbon atom at the end of a hydrocarbon chain

tertiary structure
the structure formed when covalent bonds, hydrogen bonds, ionic interactions or dispersion forces form between side groups of amino acids along a single polypeptide chain; forms the three-dimensional structure of an enzyme

theoretical yield
the amount of products expected to form from a reaction, calculated using stoichiometry

thermochemical equation
a chemical equation that shows the enthalpy change of a reaction

titrant
(in titration) the solution in the burette

titration
the addition of a standard solution to a solution of unknown concentration until the reactants have reacted in equal mole amounts

titre
the volume of solution dispensed from the burette

transesterification
the conversion of one fatty acid ester into another fatty acid ester

transition state
an intermediate state in which atoms from the reactant molecules are loosely bonded together, unstable and can spontaneously react to form new products

transmittance
the amount of infrared radiation passing through a sample

triglyceride
large non-polar molecule made from three fatty acids and a glycerol molecule

true value
the value you would get if data could be measured and recorded perfectly

U

unsaturated (molecule)
a molecule that contains at least one double or triple bond

V

valence electron
an electron in the outermost shell of an atom

valence shell
the outer shell of an atom where electrons are found

valence shell electron pair repulsion (VSEPR) theory
a theory that explains how molecules are shaped due to electrons repelling each other

validity
whether the experiment investigates what it aims or claims to investigate

vaporisation
the transition of a substance from liquid to gaseous state

viscosity
the ability of a substance to flow

volumetric analysis
a quantitative analytical technique where the volume and concentration of a solution are accurately known and are used to determine the concentration of an analyte with an accurately known volume

volumetric flask
a piece of glassware used to create a standard solution due to its ability to measure highly accurate volumes

W

water splitting
an alkaline electrolysis reaction whereby water is split into oxygen and hydrogen gases

wavenumber
a unit of frequency (cm^{-1})

Y

yield
the total amount of product formed from a chemical reaction

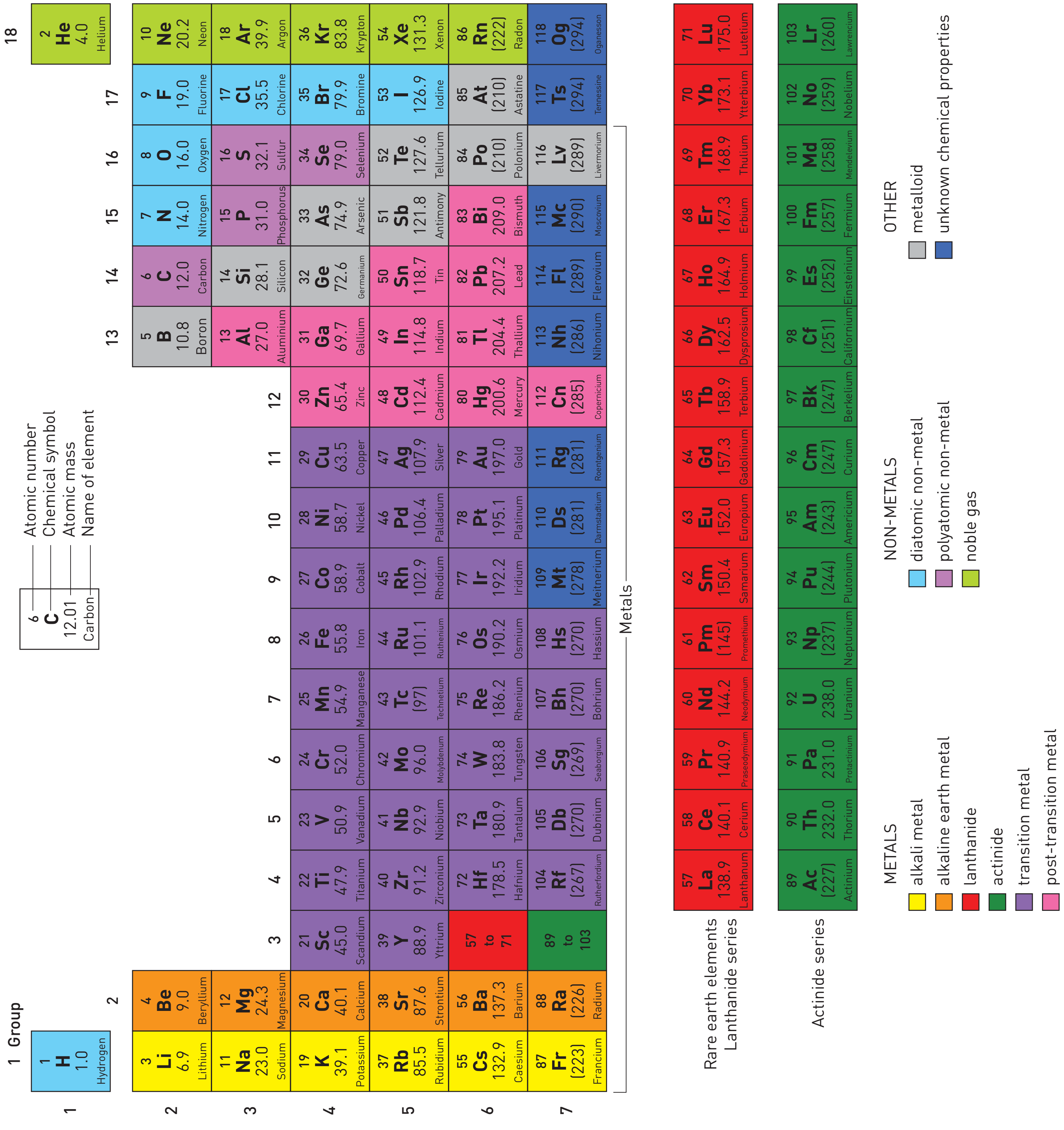
Z

zwitterion
a molecule with a positively charged group and a negatively charged group

INDEX

APPENDIX

Periodic table



6 **C** 12.01 Carbon	
6	Atomic number
C	Chemical symbol
12.01	Atomic mass
Carbon	Name of element

	1 Group	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 1.0 Hydrogen																	2 **He** 4.0 Helium
2	3 **Li** 6.9 Lithium	4 **Be** 9.0 Beryllium											5 **B** 10.8 Boron	6 **C** 12.0 Carbon	7 **N** 14.0 Nitrogen	8 **O** 16.0 Oxygen	9 **F** 19.0 Fluorine	10 **Ne** 20.2 Neon
3	11 **Na** 23.0 Sodium	12 **Mg** 24.3 Magnesium											13 **Al** 27.0 Aluminium	14 **Si** 28.1 Silicon	15 **P** 31.0 Phosphorus	16 **S** 32.1 Sulfur	17 **Cl** 35.5 Chlorine	18 **Ar** 39.9 Argon
4	19 **K** 39.1 Potassium	20 **Ca** 40.1 Calcium	21 **Sc** 45.0 Scandium	22 **Ti** 47.9 Titanium	23 **V** 50.9 Vanadium	24 **Cr** 52.0 Chromium	25 **Mn** 54.9 Manganese	26 **Fe** 55.8 Iron	27 **Co** 58.9 Cobalt	28 **Ni** 58.7 Nickel	29 **Cu** 63.5 Copper	30 **Zn** 65.4 Zinc	31 **Ga** 69.7 Gallium	32 **Ge** 72.6 Germanium	33 **As** 74.9 Arsenic	34 **Se** 79.0 Selenium	35 **Br** 79.9 Bromine	36 **Kr** 83.8 Krypton
5	37 **Rb** 85.5 Rubidium	38 **Sr** 87.6 Strontium	39 **Y** 88.9 Yttrium	40 **Zr** 91.2 Zirconium	41 **Nb** 92.9 Niobium	42 **Mo** 96.0 Molybdenum	43 **Tc** (97) Technetium	44 **Ru** 101.1 Ruthenium	45 **Rh** 102.9 Rhodium	46 **Pd** 106.4 Palladium	47 **Ag** 107.9 Silver	48 **Cd** 112.4 Cadmium	49 **In** 114.8 Indium	50 **Sn** 118.7 Tin	51 **Sb** 121.8 Antimony	52 **Te** 127.6 Tellurium	53 **I** 126.9 Iodine	54 **Xe** 131.3 Xenon
6	55 **Cs** 132.9 Caesium	56 **Ba** 137.3 Barium	57 to 71	72 **Hf** 178.5 Hafnium	73 **Ta** 180.9 Tantalum	74 **W** 183.8 Tungsten	75 **Re** 186.2 Rhenium	76 **Os** 190.2 Osmium	77 **Ir** 192.2 Iridium	78 **Pt** 195.1 Platinum	79 **Au** 197.0 Gold	80 **Hg** 200.6 Mercury	81 **Tl** 204.4 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 209.0 Bismuth	84 **Po** (210) Polonium	85 **At** (210) Astatine	86 **Rn** (222) Radon
7	87 **Fr** (223) Francium	88 **Ra** (226) Radium	89 to 103	104 **Rf** (267) Rutherfordium	105 **Db** (270) Dubnium	106 **Sg** (269) Seaborgium	107 **Bh** (270) Bohrium	108 **Hs** (270) Hassium	109 **Mt** (278) Meitnerium	110 **Ds** (281) Darmstadtium	111 **Rg** (281) Roentgenium	112 **Cn** (285) Copernicium	113 **Nh** (286) Nihonium	114 **Fl** (289) Flerovium	115 **Mc** (290) Moscovium	116 **Lv** (289) Livermorium	117 **Ts** (294) Tennessine	118 **Og** (294) Oganesson

Metals

Rare earth elements Lanthanide series	57 **La** 138.9 Lanthanum	58 **Ce** 140.1 Cerium	59 **Pr** 140.9 Praseodymium	60 **Nd** 144.2 Neodymium	61 **Pm** (145) Promethium	62 **Sm** 150.4 Samarium	63 **Eu** 152.0 Europium	64 **Gd** 157.3 Gadolinium	65 **Tb** 158.9 Terbium	66 **Dy** 162.5 Dysprosium	67 **Ho** 164.9 Holmium	68 **Er** 167.3 Erbium	69 **Tm** 168.9 Thulium	70 **Yb** 173.1 Ytterbium	71 **Lu** 175.0 Lutetium
Actinide series	89 **Ac** (227) Actinium	90 **Th** 232.0 Thorium	91 **Pa** 231.0 Protactinium	92 **U** 238.0 Uranium	93 **Np** (237) Neptunium	94 **Pu** (244) Plutonium	95 **Am** (243) Americium	96 **Cm** (247) Curium	97 **Bk** (247) Berkelium	98 **Cf** (251) Californium	99 **Es** (252) Einsteinium	100 **Fm** (257) Fermium	101 **Md** (258) Mendelevium	102 **No** (259) Nobelium	103 **Lr** (260) Lawrencium

METALS
- ■ alkali metal
- ■ alkaline earth metal
- ■ lanthanide
- ■ actinide
- ■ transition metal
- ■ post-transition metal

NON-METALS
- ■ diatomic non-metal
- ■ polyatomic non-metal
- ■ noble gas

OTHER
- ■ metalloid
- ■ unknown chemical properties

This front cover shows glow sticks emitting light. When a glow stick is activated, hydrogen peroxide reacts with oxalate esters to form an intermediate. This intermediate then reacts with, and excites, a coloured dye. As the dye relaxes, it releases coloured light.

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